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Electrochemical Changes

Unit B

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Chemistry 30

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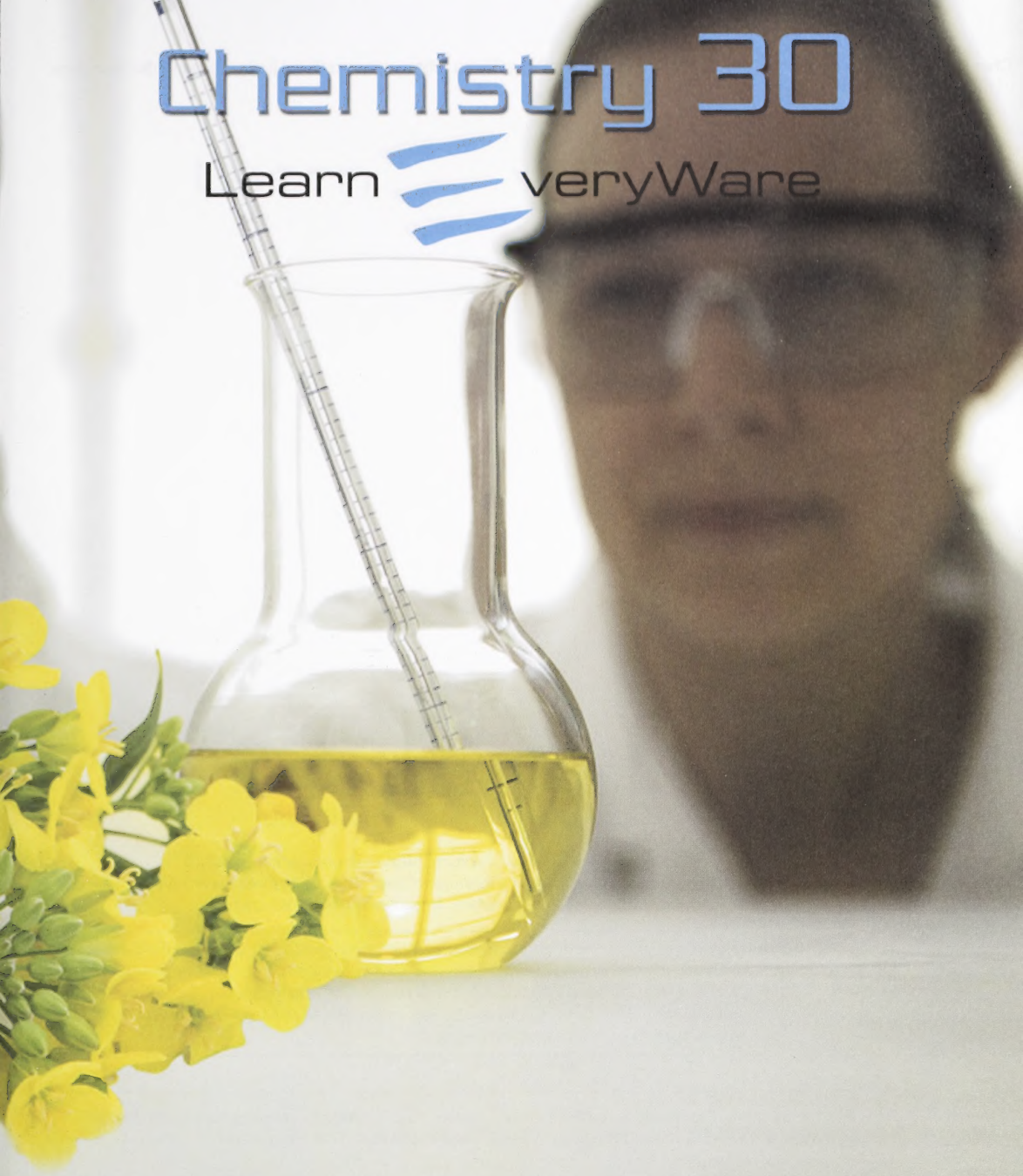

Electrochemical Changes

Unit B

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- Calgary Board of Education
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Unit B: Electrochemical Changes
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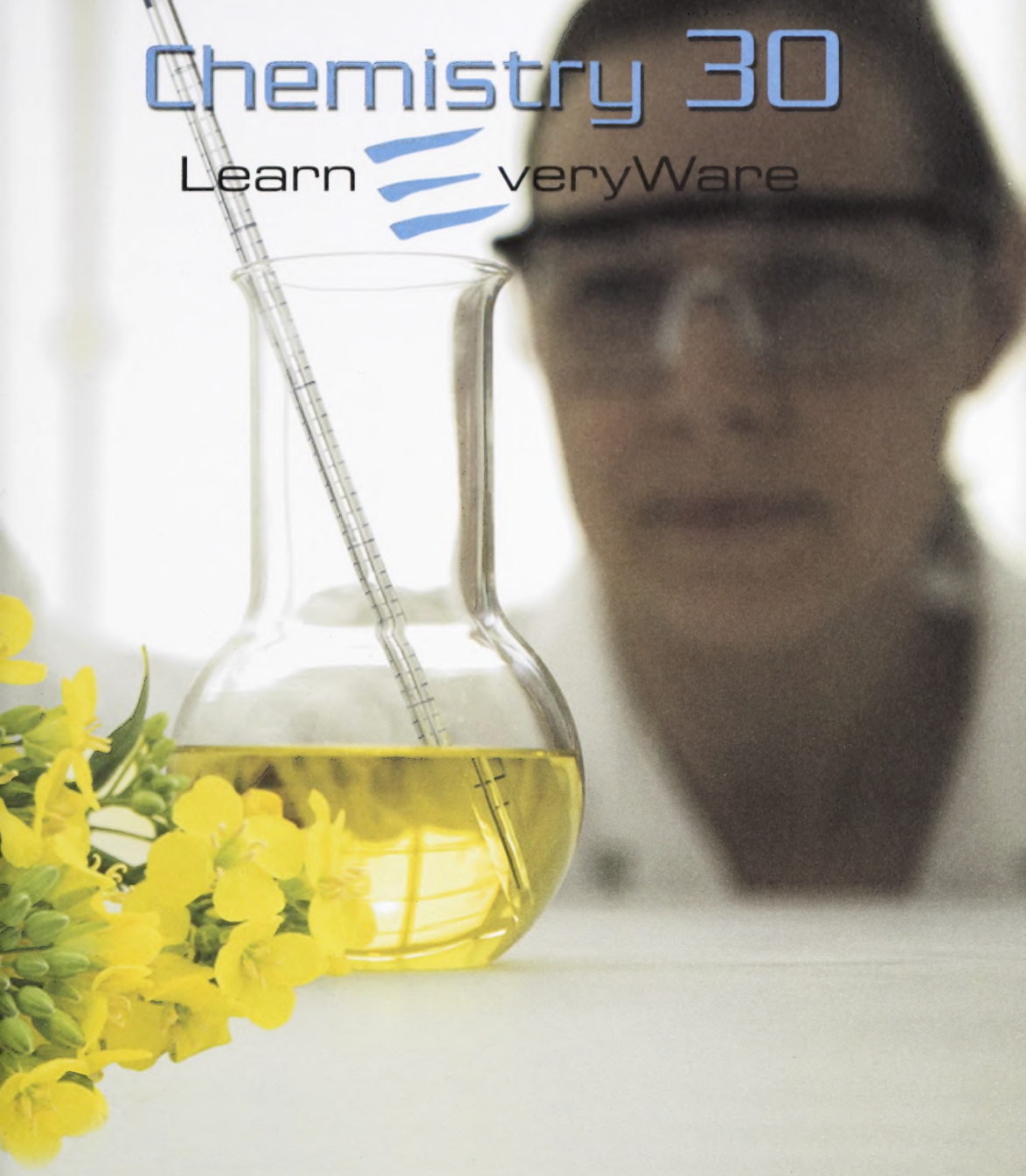
Chemistry 30

Learn EveryWare



Electrochemical Reactions

Module 3

Contents

Unit B Introduction 2

Concept Organizer 3

Module 3 Introduction 6

Big Picture 7

In This Module 8

Lesson 1: Electrochemical Change 10

Lesson 2: Predicting Redox Reactions 15

Lesson 3: Half-Reactions 19

Lesson 4: Oxidation Numbers and Corrosion Protection 25

Lesson 5: Redox Stoichiometry 32

Module Summary 37

Module Assessment 38

Module Glossary 40

Appendix (Self-Check Answers) 41

Unit B Introduction

Unit B focuses on electrochemical changes in chemical systems. You will explore these concepts by learning the principles involved in the operation of a technology you probably use every day—a commercial electric cell, sometimes called a battery.

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In this unit you will look at chemical systems, and you will investigate them as electrochemical systems—systems involving an exchange of electrons. Electrochemical systems are numerous and varied. Electrochemical systems include, for example, biological and non-biological systems and natural and technological systems.

Module 3 introduces the scientific principles associated with reduction-oxidation (redox) reactions. You will learn how redox reactions can be used to analyze and predict changes in a chemical system and to perform quantitative analysis. This module will support the work you will do in Module 4.

Module 4 explores how technologies, like the commercial electric cell, involve an application of redox reactions. You will use your knowledge of electrochemical change to observe, analyze, and evaluate the design and function of electric and electrolytic cells.

Together these modules develop your understanding that some types of chemical change involve a transfer of electrons. In the last unit of this course you will explore chemical reactions that involve proton transfer reactions. You will apply many of the principles you learn in Unit B to your work throughout this course.

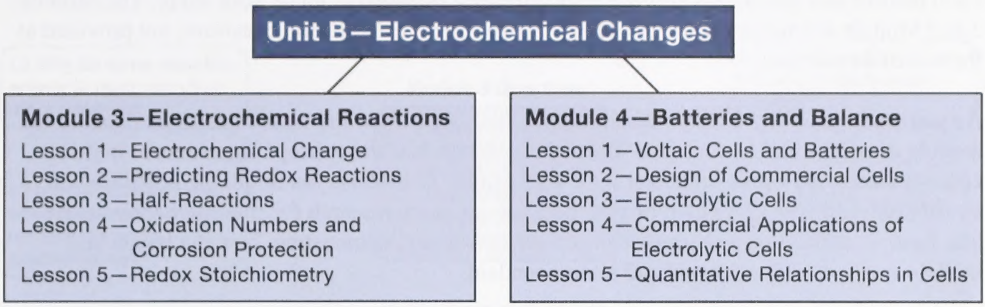
At the end of Unit B you will be able to

- explain the nature of oxidation-reduction reactions
- apply the principles of oxidation-reduction to electrochemical cells
- analyze scientific and technological systems that apply electrochemical principles

In the Unit B Assessment you will explore the scientific and technological aspects of reduction-oxidation reactions, and you will explain and identify their positive and negative characteristics.

Concept Organizer

Unit B—Electrochemical Changes


Unit B—Electrochemical Changes
Module 3—Electrochemical Reactions
Lesson 1—Electrochemical Change
Lesson 2—Predicting Redox Reactions
Lesson 3—Half-Reactions
Lesson 4—Oxidation Numbers and Corrosion Protection
Lesson 5—Redox Stoichiometry
Module 4—Batteries and Balance
Lesson 1—Voltaic Cells and Batteries
Lesson 2—Design of Commercial Cells
Lesson 3—Electrolytic Cells
Lesson 4—Commercial Applications of Electrolytic Cells
Lesson 5—Quantitative Relationships in Cells

Module Descriptions

Module 3—Electrochemical Reactions

Module 3 introduces the scientific principles associated with reduction-oxidation reactions. You will learn how redox reactions can be used to analyze and predict changes in a chemical system, and you will learn how to perform quantitative analysis. This module will support the work you will do in Module 4.

You will consider the following module questions:

- What properties of metals make them popular choices in the construction and production of materials?
- How can an understanding of corrosion allow for better selection of materials and development of methods that reduce material damage?

Module 4—Batteries and Balance

Module 4 explores how technologies, like the commercial electric cells you use to operate your cell phone or other electronic devices, involve an application of redox reactions. You will use your knowledge of electrochemical change to observe, analyze, and evaluate the design and function of electric and electrolytic cells.

You will consider the following module question:

- How are principles of oxidation and reduction applied to electrochemical cells?

Together these modules develop your understanding that some types of chemical change involve a transfer of electrons. In the last unit of this course you will explore chemical reactions that involve proton transfer reactions. Many of the principles you will learn in Unit B will apply to your work throughout this course.

Using the Concept Organizer

Each module and lesson contains focusing questions intended to guide your study. The Module 3 and Module 4 concept organizers, which list the module and lesson questions, are provided at the end of this section.

As you work your way through each module, think about how the lesson questions relate to the module questions and to questions from other lessons. Use the concept organizers to build a concept map or a graphic organizer for each module. To do this, you might use software you already have on your computer, or you might do an Internet search for free software you can use. In your concept map or graphic organizer, record and demonstrate how the lesson and module concepts are connected and interdependent.

Sample concept maps are provided in the Module Summary for each module. Remember that the samples are just that—they show only one of many possible descriptions. However, if your completed concept map or graphic organizer differs significantly from the sample, you may wish to contact your teacher or to compare your map or organizer to those of your classmates. This will ensure that your interpretations of lesson materials and your descriptions are accurate.

Module 3 Concept Organizer

L1 What is an electrochemical change?

L2 Why do some materials appear to react more than other materials?

L2 Is it possible to predict whether an electrochemical reaction will occur spontaneously?

L3 How can combinations of species act together as oxidizing or reducing agents?

L3 Can half-reactions be used to predict and explain changes that occur within a chemical system?

Module 3 Questions

What properties of metals make them popular choices in the construction and production of materials?

How can an understanding of corrosion allow for better selection of materials and development of methods that reduce material damage?

L3 Can the same substance be the oxidizing agent and the reducing agent in an electrochemical process?

L5 How is the stoichiometric method applied to redox systems?

L5 How can a chemical system be analyzed using redox reactions?

L4 Can corrosion be prevented?

L4 What factors cause corrosion?

L4 What are oxidation numbers and how can they be used to understand redox reactions?

Module 4 Concept Organizer

L1 How much do you rely on batteries in your everyday life?

L1 What are the components of a voltaic cell?

L2 Why are there so many different designs and types of electrochemical cells?

Module 4 Questions

How are principals of oxidation and reduction applied to electrochemical cells?

L3 What is an electrolytic cell?

L5 What are the important quantitative relationships within operating voltaic and electrolytic cells?

L4 What are some of the practical applications of electrolytic cells?

L3 What chemical and energy conversions occur within an electrolytic cell?

Module 3—Electrochemical Reactions

Module Introduction

Metals are fantastic materials to use for structural components or for decoration. Sometimes, however, metals react when they are exposed to air, water, or other substances. In some situations the reaction decreases the performance of the object. In extreme cases, metal fatigue or failure can have drastic consequences, such as when metal fatigue causes a large structure to collapse.

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A great deal of scientific work and technological development has gone into studying metals and into preventing metal corrosion. If you consider how many of the objects you use involve metals, you will understand why so much attention is placed on understanding corrosion.

In Module 3 you will investigate the following questions:

- What properties of metals make them popular choices in the construction and production of materials?
- How can an understanding of corrosion allow for better selection of materials and for development of methods that reduce material damage?

Remember that each lesson will also be organized around questions intended to guide your study. As you proceed through Module 3, you may record answers to these questions and any interrelationships that exist between them in a concept map or graphic organizer. More information is available in the Unit B Concept Organizer. In the Module 3 Summary you will receive further information on how you can use your concept map or graphic organizer to review the concepts you studied in this module.

Big Picture

In 1995, the High Level Bridge in Edmonton, Alberta, underwent a major restoration in which approximately 45 tonnes of rusted steel were replaced, and 100 000 litres of paint were applied. How could so much rust form? Why was paint applied at the end of the restoration?

Look around and you will see many metal objects. Metals are heavily used by society because of their properties. For example, metals are electrical conductors, are strong, and can be made into different shapes.

One negative property of metals is their reactivity. The performance of metals can be negatively affected when metals undergo chemical change. Metal performance is an issue in large structures such as bridges and in small devices that use metals to conduct electric current.

You will learn many theoretical concepts in Module 3. It is important that you identify where you are able to use these theoretical concepts to predict and explain your observations and where you are not able to do so.

In Module 4 you will investigate many technological applications of the main theoretical concept of Module 3—reduction-oxidation reactions. Successfully applying your theoretical knowledge to new situations will help you in Module 4.

Assessment in This Module

Each lesson contains a range of activities and assessment options. These include assignments, labs, and Self-Check, Try This, Discuss, Reflect and Connect, and Reflect on the Big Picture activities. Instructions will be provided for each of these activities so that you can appropriately focus your time and effort. Your teacher will tell you which assessment options to complete and which responses to submit for marks or feedback. Remember to save all of your work in your Chemistry 30 folder.

In the Module 3 Assessment you will apply your knowledge of electrochemical change to describe the electrochemical basis behind methods of preventing corrosion.

You may wish to look at the Module Assessment and the Unit Assessment before starting Lesson 1.

In This Module

Lesson 1—Electrochemical Change

In Lesson 1 you will observe some chemical reactions in a new way—as electrochemical changes. You will learn to write chemical reactions that demonstrate the kind of change that occurs in these systems, and you will learn how to interpret these reactions.

You will investigate the following lesson question:

- What is an electrochemical change?

Lesson 2—Predicting Redox Reactions

In Lesson 2 you will analyze sets of data collected from the reaction of metals with aqueous metal ions to identify patterns in the reactivity of metals and their ions. You will use these patterns to predict reactions involving other substances.

You will investigate the following lesson questions:

- Why do some metals appear to react more easily than other metals?
- Is it possible to predict whether an electrochemical reaction will occur spontaneously?

Lesson 3—Half-Reactions

In Lesson 3 you will learn to write and use half-reactions involving more than one oxidizing or reducing agent. You will use half-reactions and net ionic equations to explain chemical change in systems in which the half-reactions for the oxidizing and reducing agents are not provided to you.

You will investigate the following lesson questions:

- How can combinations of species act together as oxidizing or reducing agents?
- Can half-reactions be used to predict and explain changes that occur within a chemical system?
- Can the same substance be the oxidizing agent and the reducing agent in an electrochemical process?

Lesson 4—Oxidation Numbers and Corrosion Protection

In Lesson 4 you will interpret empirical changes in terms of oxidation states (sometimes as distinctive as colour changes by certain metal ions) or numbers used to explain the oxidation or reduction of atoms.

You will investigate the following lesson questions:

- What are oxidation numbers, and how can they be used to understand redox reactions?
- What factors cause corrosion?
- How can corrosion be prevented?

Lesson 5—Redox Stoichiometry

In Lesson 5 you will learn how to adapt your knowledge of redox reactions into designing and performing quantitative analysis. You will investigate quantitative relationships using a titration experiment involving oxidizing and reducing agents. In Module 4 you will extend this to understanding other quantitative measures.

You will investigate the following lesson questions:

- How can a chemical system be analyzed using redox reactions?
- How is the stoichiometric method applied to redox systems?

Module 3—Electrochemical Reactions

Lesson 1—Electrochemical Change

Get Focused

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Corrosion is a common sight on metal objects. Often we think of corrosion as the rusting of iron, but it can also be used to describe the change in other metals, like the copper in the statue shown in the picture.

What colour does the metal copper have? Can you explain what has happened to the copper in the statue? How could the copper metal change into copper ions, which appear to wash away from the statue? Is this an example of a chemical reaction and, if so, what happens when atoms change in this manner?

The type of change observed here, in which atoms appear to change form and charge, is an important type of change called electrochemical change. Electrochemical change is the focus of this unit.

Consider the following question as you complete Lesson 1:

- What is an electrochemical change?

Module 1: Lesson 1 Assignment

There is no assignment for this lesson.

There are other questions in this lesson that are not marked by the teacher; however, you should still answer these questions. The Self-Check, Try This, and other types of questions are placed in this lesson to help you review important information and build key concepts that may be applied in future lessons. You should record the answers to all the questions in the lesson and place those answers in your course folder.

After a discussion with your teacher, you must decide what to do with the questions that are not part of your assignment. For example, you may decide to submit the responses to Try This and other questions that are not marked to your teacher for informal assessment and feedback. Your answers are very important to your teacher. They provide your teacher with information about your learning, and they help your teacher identify where adjustments to your instruction may be necessary.

Explore

Try This

Have you ever noticed when you look at the jar of spare change in your house that pennies really lose their shine? Is there a relationship between the age of the pennies and their loss of luster (shininess)? Look at the dates on some of the pennies in your household change jar or in your pocket to see if there is a connection between a penny's age and its luster.

TR 1. Write a hypothesis describing a relationship between a possible factor and the loss of luster by copper as demonstrated by the pennies in your household change jar. Submit a copy of your hypothesis to your teacher for feedback.

Read

Read page 556 of the textbook.

Do you think the change in copper observed in the pennies or in the photograph of the copper statue at the beginning of this lesson demonstrates an electrochemical process? Start your analysis of these systems by writing the chemical formula for the forms of copper that have the appearances you observed. Can you describe the differences between the chemical formulae for these forms of copper? Do the differences in the chemical formulae you have written fit the description of an electrochemical process (a type of chemical change that involves a transfer of electrons)?

Read

In previous science courses you learned that the net charge given to an atom or group of atoms is due to the quantity of protons and electrons the atom or group possesses. Electrochemical changes involve a transfer of electrons. Therefore, a change in the net charge of a substance often indicates that the substance has undergone an electrochemical change.

In the case of copper, its various forms have distinct colours, and the change in copper from its elemental form (net charge zero) to copper(II) ion (net charge 2+) involves a loss of electrons to create the positively charged copper(II) ion.

Some of the first investigations involving electrochemical reactions occurred as society began to work with metals. Read pages 558–559 of the textbook to see how metallurgy developed some of the terminology you will use throughout this unit.

Lesson 1 Lab: Observing the Reactivity of Zinc

What happens when a metal is involved in a chemical reaction?

Go to the course multimedia DVD, and view the virtual investigation "Observing the Reactivity of Zinc." You will observe the changes that occur when zinc reacts with a solution of silver nitrate. Record your observations and answer the Analysis questions as you work through this investigation. Save a copy of your work in your course folder.

Analysis

1. Write the balanced chemical equation for the reaction of zinc and aqueous silver nitrate.

2. Indicate what empirical evidence you observed in the investigation that supports the equation you have written.

3. Identify which substances have undergone change in this reaction? Is the change the substance(s) underwent an electrochemical change involving a transfer of electrons? Support your reasoning.

In the next activity, you will read and work through "Sample problem 13.1" in the textbook. You will be prompted to check your answers to the Analysis Questions after completing your reading.

Read

Electrochemical change involves a transfer of electrons. In the virtual investigation, "Observing the Reactivity of Zinc," you saw that silver ions must have gained electrons to become solid silver. How is such a change represented?

Read pages 561–564 of the textbook to learn how to represent electrochemical changes using a balanced chemical equation. Carefully work through "Sample problem 13.1" and "Communication example 1" on page 563 to practise using chemical equations to explain electrochemical changes.

Remember to use this reading (particularly "Sample problem 13.1") to check your answers to the analysis questions for the virtual investigation, "Observing the Reactivity of Zinc." If you identify significant differences between your answers and the textbook reading, you may wish to contact your teacher for help.

Half-Reactions

In the textbook reading you will have noticed that half-reactions are used to describe the changes observed in chemical systems. Half-reactions are one type of balanced chemical equation. You will recall from previous science courses that balanced chemical equations demonstrate two conservation laws:

- law of conservation of mass—in any physical or chemical change, the total initial mass of reactant(s) is equal to the total final mass of product(s)
- law of conservation of charges—in any isolated system, the sum of all of the charges remains constant

Conservation of mass is quite easy to identify in an equation because the type and numbers of each type of atom are the same on both sides of the equation. For example,

$$Ag^{+}(aq) + 1\ e^{-} \rightarrow Ag(s)$$

Reactants	Products
Ag = 1	Ag = 1

Conservation of charge involves adding up the charges on each side of the equation. If the net charge on each side is equal, the equation is balanced for charge. For example,

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^{-}$$

Reactants	Products
Cu = 1	Cu = 1
charge = 0	charge = +2 + (–2) = 0

Self-Check

SC 1. Complete "Practice" questions 7–11 on page 564 of the textbook.

SC 2. Use the table "Colours of Common Aqueous Ions" on page 11 of the Chemistry 30 Data Booklet or a similar table to determine the colour of the resulting solution if nitric acid were reacted with each of the following metals:

a. nickel
b. copper
c. chromium

SC 3. Write balanced half-reactions for the oxidation and reduction that occurred in the virtual investigation "Observing the Reactivity of Zinc."

Check your work with the answer in the appendix.

Reflect on the Big Picture

Understanding electrochemical reactions is an important part of successfully using metals as structural materials. In this lesson you have seen how metals can undergo electron transfer reactions.

RBP 1. How do metals tend to behave in electrochemical reactions? What possible consequence might a redox reaction involving a metal have to the substance's integrity? Can you think of ways that electrochemical reactions of metals can be prevented?

Save your response in your course folder.

Lesson Summary

In Lesson 1 you considered the following question:

- What is an electrochemical change?

By observing the reaction between zinc and aqueous silver ions you were able to see evidence of a chemical reaction that can be explained by an electron transfer between reactants. You also learned how to write, balance, and interpret half-reactions that describe reduction or oxidation processes.

In the next lesson you will investigate electrochemical changes involving other metals, and you will conduct an experiment to investigate the difference in the reactivity of selected metals.

Module 3—Electrochemical Reactions

Lesson 2—Predicting Redox Reactions

Get Focused

Each metal differs in its ability to react. For example, when exposed to air, pennies (primarily made of copper) tend to lose their luster. Silver coins and jewelry tend to tarnish and turn black. Gold, however, does not appear to lose its luster when exposed to air.

Rust on a household strainer. Was the selection of iron a good choice of metal for the strainer considering the intended use for this object?

In Lesson 1 you learned that the loss of luster in some metals is due to an electrochemical reaction in which the metals are oxidized and exchange electrons with another reactant. Knowledge of which metals have the greatest potential to oxidize, and in what conditions their oxidation is rooted, could be a great asset when considering materials for use in construction or other applications.

Consider the following questions as you complete Lesson 2:

- Why do some metals appear to react more easily than other metals?
- Is it possible to predict whether an electrochemical reaction will occur spontaneously?

The focus of Lesson 2 is on learning how to interpret reaction data that you will collect. You will also learn to construct tables that will help you make predictions about the reactivity of metals or metal ions with other substances.

Being able to interpret reaction data in the manner you will learn in this lesson is essential to your study in the rest of this unit—be extra careful as you complete questions and work through problems in this lesson, and ensure you have mastered this skill before moving on to the next lesson.

Module 3: Lesson 2 Assignment

You will be directed to go to the Assignment Booklet to complete the Module 3: Lesson 2 Assignment later in this lesson. The assignment has the following parts:

- Part 1: Lab—Spontaneity of Redox Reactions
- Part 2: Lab Exercise 13.A on page 572 of the textbook

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Read

The relative reactivity or stability of a metal is determined empirically. Later in this lesson you will complete an investigation in which you will observe and interpret the data from many systems of metals and aqueous solutions of metal ions.

To prepare for the investigation, read page 568 and the top half of page 569 in the textbook. You will learn about the terms *oxidizing agent* and *reducing agent*. These terms are commonly used to discuss reactants involved in an electrochemical reaction.

Evidence of a chemical reaction can take many forms. List observations that indicate that a chemical reaction has taken place.

When a colour change, evolution of a gas, formation of a precipitate, or temperature change of a system is observed, the reaction can be described as **spontaneous**; that is, the chemical change occurs without a continuous input of energy. If no reaction evidence is observed, the reaction is **non-spontaneous**.

spontaneous: a process that occurs without any external energy source

A spontaneous process will occur on its own.

non-spontaneous: a process that is incapable of proceeding unless driven by an outside source of energy

In the next investigation you will determine whether the reaction between various metals and solutions of metal ions is spontaneous or non-spontaneous.

Lesson 2 Lab: Spontaneity of Redox Reactions

Read "Investigation 13.2" on page 602 in the textbook. In the virtual investigation you will complete, lead and the solution of lead(II) ion were removed. Exposure to lead and lead compounds is a concern. You may wish to consult an MSDS (Materials Safety Data Sheet) or other reference on safety in science laboratories for more information about the safety concerns associated with lead or other substances you may encounter in this course.

Open the Assignment Booklet. You will record data and complete your analysis in "Part 1: Lab—Spontaneity of Redox Reactions" of the "Module 3: Lesson 2 Assignment."

Go to the course multimedia DVD, and view the virtual investigation "Spontaneity of Redox Reactions."

Module 3: Lesson 2 Assignment

Open the Assignment Booklet to the "Module 3: Lesson 2 Assignment." Complete "Part 1." You will receive information later in the lesson on when to complete the rest of the Lesson 2 Assignment.

Read

Metals that are more likely to react with the ions of other metals are considered to be more reactive or, using the terminology introduced in this lesson, are stronger reducing agents. A redox table is used to list oxidizing and reducing agents by their strengths. Read the lower half of page 569 to the end of the "Summary" section on page 571 of the textbook to learn more about the organization of a redox table.

Self-Check

SC 1. Complete "Practice" questions 1–3 and 5–10 on page 571 of the textbook.

Check your work with the answer in the appendix.

Read

In SC 1 you were asked to develop a hypothesis about the design of a redox table and to test the redox table using the data available to you.

Read pages 572–574 of the textbook. Carefully work through "Sample problem 13.4" on page 573 and "Figure 6" in the margin of page 573 to understand the spontaneity rule, a general principle demonstrated by redox tables constructed in the manner described in this lesson. You may wish to refer to page 7 of the Chemistry Data Booklet on which an extended redox table is shown.

Self-Check

SC 2. Complete "Practice" questions 11–14 on page 573 of the textbook.

SC 3. Complete "Practice" question 20 on page 574 of the textbook.

Check your work with the answer in the appendix.

Module 3: Lesson 2 Assignment

Go to the Assignment Booklet, and complete "Part 2" of the "Module 3: Lesson 2 Assignment."

Reflect and Connect

RC 1. In the Get Focused section at the beginning of this lesson, three commonly used metals were identified: copper, silver, and gold. Examine the position of these metals on the large table of half-reactions shown in your Chemistry Data Booklet. How do the locations of these metals in the table correlate with the observations made in the Get Focused section?

Save your response in your course folder.

Lesson Summary

In Lesson 2 you considered the following questions:

- Why do some metals appear to react more easily than other metals?
- Is it possible to predict whether an electrochemical reaction will occur spontaneously?

In this lesson you analyzed data collected from the reaction of metals with solutions of metal ions, and you developed a list ranking metals and metal ions in terms of their tendency to react. You also learned how to classify substances as oxidizing and reducing agents, and you learned how to rank them in terms of their tendency to transfer electrons. You learned about the spontaneity rule and how to use a table of redox reactions to predict the outcome of electrochemical reactions.

Lesson Glossary

non-spontaneous: a process that is incapable of proceeding unless driven by an outside source of energy

spontaneous: a process that occurs without any external energy source

A spontaneous process will occur on its own.

Module 3—Electrochemical Reactions

Lesson 3—Half-Reactions

Get Focused

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Electrochemical reactions involve a transfer of electrons between species. So far in this module you have learned to identify oxidation, reduction, and agents that bring about the oxidation and reduction of other substances. In Lesson 2 you wrote half-reactions for simple changes to the charge of metals and metal ions. You also used empirical evidence, collected as you completed the investigation "Spontaneity of Redox Reactions," to confirm the half-reactions you prepared.

You also began using the half-reactions listed on the "Table of Selected Standard Electrode Potentials" on page 7 of the Chemistry Data Booklet. Did you look at some of the other reactions in the table? If so, you may have noticed that many of the half-reactions involve more than one species.

Understanding how substances act alone or in combinations to bring about reactions in a chemical system is an important part of designing chemical systems to achieve a desired function. Chemists and technologists are required to investigate the behaviour of materials to ensure the materials will not undergo unwanted side reactions.

In some situations unexpected side reactions can occur when air, water, or other substances are present in a system. If their presence was not expected, an electrochemical reaction could result in the metal's corrosion, which could lead to metal fatigue and/or to a change in the metal's ability to function as intended.

Consider the following questions as you complete Lesson 3:

- How can combinations of species act together as oxidizing or reducing agents?
- Can half-reactions be used to predict and explain changes that occur within a chemical system?
- Can the same substance be the oxidizing agent and the reducing agent in an electrochemical process?

Module 3: Lesson 3 Assignment

You will be directed to go to the Assignment Booklet to complete the Module 3: Lesson 3 Assignment later in this lesson. In this assignment you will develop an experimental design and complete the data collection and analysis for the lab "Testing Predictions."

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Read

In Lesson 1 you learned that balancing chemical reactions involves ensuring that charge is balanced. When writing half-reactions, you will have checked that the net charge on each side of the equation was equal. In more complex half-reactions, balancing charge requires more care.

Read the section "Writing Complex Half-Reaction Equations" and work through the "Sample problems" and "Communication examples" on pages 564–567 of the textbook. You may wish to place a copy of "Summary" on page 567 in your course folder to review later in this module.

Self-Check

SC 1. Complete "Practice" question 12 on page 566 of the textbook.

Check your work with the answer in the appendix.

Read

When selecting materials for use in devices, it is useful to know what substances are present in the environment in which the device will be used. For example, if design manufacturers are considering using metals in machinery that will be exposed to water and oxygen, they need to know that those conditions may promote corrosion of the metal.

Look at the list of half-reactions in your Chemistry Data Booklet. Can you spot the half-reactions for iron and other metals? Can you locate the half-reaction for oxygen in combination with water? Is the position of the half-reaction containing oxygen and water acting as an oxidizing agent in a position that would result in a spontaneous reaction with iron?

Read page 575 to the end of "Communication example 1" on page 577 of the textbook. A summary of the method described in this section appears on page 578.

Self-Check

SC 2. Complete "Practice" question 26 on page 579 of the textbook.

Check your work with the answer in the appendix.

Lesson 3 Lab: Testing Predictions

Background Information

In this lesson you learned to predict redox reactions using the five-step method listed on page 578 of the textbook. Are the predictions made using this method supported by empirical evidence? Having a method to predict and explain changes in metals, metal ions, or other substances in chemical systems would have great value when considering chemical systems that might exist beyond the ones you construct in a laboratory.

For instance, consider the equipment used in the extraction of bitumen in Alberta's oil sands operation. Electrochemical changes may be suspected in cases of metal fatigue in the equipment. Could the five-step method be used to identify alternative metals for use in this equipment?

In the Pre-Lab for this investigation you will design an experiment to test predictions using the following combinations of reactants:

System	Reactant 1	Reactant 2
1	HCl(aq)	Cu(s)
2	HCl(aq)	Mg(s)
3	$Cu(NO_3)_2(aq)$	Sn(s)
4	NaOH(aq)	$I_2(aq)$
5	$Cu(NO_3)_2(aq)$	Mg(s)

As you design your experiment, make sure you consider safety aspects when suggesting how these substances should be manipulated and how waste from the reactions should be disposed of.

As you saw in SC 2, it is important that you be able to use diagnostic tests, detect colour changes, measure changes to pH, or identify other evidence of reaction. The evidence of reaction that you predict and make certain you observe will allow you to form a conclusion about the method you are testing in this investigation.

Go to the Assignment Booklet. Complete the "Pre-Lab" section of the "Module 3: Lesson 3 Assignment" now. You will record your observations in your Assignment document as you view the virtual investigation, and then you will complete the "Analysis" section.

Go to the course multimedia DVD, and view the virtual investigation "Testing Predictions for Redox Reactions."

Module 3: Lesson 3 Assignment

If you have not already done so, open the Assignment Booklet. Complete the "Module 3: Lesson 3 Assignment."

Read

Some substances are not stable. One way to explain that lack of stability is to describe a reaction in which the same species acts as both the oxidizing and the reducing agent. Read "Disproportionation" on page 577–578 of the textbook.

Read

In this lesson you have learned how to

- use a redox table to predict the spontenaeity of reactions
- write half-reactions for species not on a redox table

Is it possible to add any oxidation and reduction half-reactions to produce a net ionic equation? Read "Predicting Redox Reactions by Constructing Half-Reactions" on pages 579–581 in the textbook. You will note that the method is very similar to the method you have already learned.

So far in this module, you have learned many skills for writing half-reactions and net ionic chemical equations, and for predicting whether spontaneous chemical change will occur when reactants are combined. You have also learned how to write chemical equations for substances that are not listed on the "Table of Selected Standard Electrode Potentials" in your Chemistry Data Booklet. If you write reactions that do not appear on this table, what aspect of this chemical system are you unable to determine?

Self-Check

SC 3. Complete "Practice" question 31 on page 581 of the textbook.

Check your work with the answer in the appendix.

Reflect on the Big Picture

Early European explorers were surprised to find that the Inuit of northwestern Greenland were hunting with iron weapons despite the fact that the Inuit did not appear to have the technology to separate iron from its ore. To add to the puzzle, at the time, Greenland had no known iron ore deposits. Essentially, the Iron Age had begun for the First Nations people of northwestern Greenland. This was interesting because the use of iron was not documented anywhere else in the western hemisphere until contact was made with Europeans.

The mystery of how the Inuit had access to iron was solved when the explorers discovered that the Inuit were harvesting the iron from several huge meteorites. The Ahnighito meteorite pictured is 31 tonnes and is the largest fragment of a 59-tonne mass called the Cape York meteorite, which is believed to be the center of an early planet. The Ahnighito meteorite and some of the other, smaller fragments are on display in the American Museum of Natural History in New York City.

Iron found in ore deposits on Earth exists in an oxidized state, for example, as Fe_2O_3. Much of the iron found in a meteorite is in the form of the element, Fe. The Inuit were able to make tools using pieces of the Cape York meteorite because the pieces contain over 90% iron. Since meteorites form in space, many meteorites contain unique minerals (iron compounds) not normally found on Earth.

One of the most studied meteorites to strike Canada is the Abee meteorite, which struck Earth in 1952 near Abee, Alberta. The Abee meteorite is only 32% iron. Accounts indicate that the Abee meteorite smells like gunpowder due to the large proportion of sulfur it contains. The unique iron and sulfur minerals in this meteorite, such as djerfisherite ($K_3CuFe_{12}S_{14}$), may suggest that the meteorite was once a piece of the planet Mercury and was formed over 4.5 billion years ago. The meteorite was extracted from a two-meter-deep hole in a wheat field by a farmer, a school teacher, and several other interested townsfolk. The Abee meteorite is currently part of the National Meteorite Collection and remains available for further study.

When *Chemistry 30 Learn EveryWare* was being developed, a meteorite was observed in the night skies over Alberta. Pieces were found near Lone Rock, Saskatchewan. You might research the composition of this meteorite and the theories about its origin.

RBP 1. Use your knowledge of electrochemistry to explain how it is possible for iron to exist in different forms (both as a pure substance and within compounds) found in Earth and in space as evidenced by the composition of the Cape York and Abee meteorites.

Save a copy of your response in your course folder.

Lesson Summary

In Lesson 3 you considered the following questions:

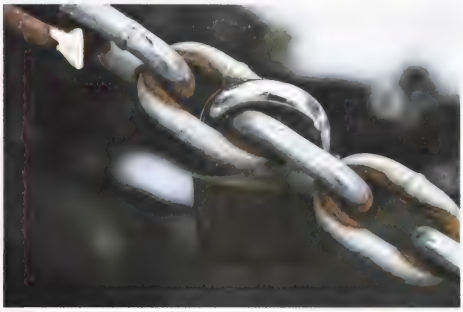
- How can combinations of species act together as oxidizing or reducing agents?
- Can half-reactions be used to predict and explain changes that occur within a chemical system?
- Can the same substance be the oxidizing agent and the reducing agent in an electrochemical process?

Redox reactions can involve a variety of substances, which sometimes act in combination to promote the oxidation or reduction of other substances. You learned that half-reactions can be used to describe the chemical changes that occur within a system. If provided with a table of half-reactions written as reductions, you are able to predict which combinations of substances will react spontaneously.

Module 3—Electrochemical Reactions

Lesson 4—Oxidation Numbers and Corrosion Protection

Get Focused



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Throughout this module you have considered the reaction of metals with other substances. You have learned many techniques for interpreting the changes observed in chemical systems undergoing electrochemical change. You have also learned many new skills, namely the ability to predict the spontaneity of reactions and to explain disproportionation reactions.

Many structures rely on iron for their integrity. For example, reinforced concrete contains rods of steel, which is primarily composed of iron, called rebar. The frames of automobiles also rely on iron for their structural integrity. But we know iron can rust and, when it does, it loses its integrity. In the Module 3 Big Picture you read about the restoration of the High Level Bridge in Edmonton, Alberta, and the quantity of corroded material that was removed from the bridge. Since so many materials rely on iron or steel, corrosion has a great economic impact and presents serious safety concerns.

In Lesson 4 you will investigate the process of corrosion and how it can be further understood as a redox process. You will also learn about methods that are used to prevent the corrosion of iron.

Consider the following questions as you complete Lesson 4:

- What are oxidation numbers and how can they be used to understand redox reactions?
- What factors cause corrosion?
- How can corrosion be prevented?

Module 3: Lesson 4 Assignment

You will be directed to go to the Assignment Booklet to complete the Module 3: Lesson 4 Assignment later in this lesson. The assignment has the following three parts:

- Part 1: Lab—Oxidation States of Manganese
- Part 2: Lab—Corrosion of Nails
- Part 3: Questions

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Try This

Iron Corrosion

The corrosion of iron is a complex electrochemical process. Steel is primarily composed of iron. Its corrosion occurs more frequently in moist conditions in which the iron is also exposed to oxygen.

The process of corrosion involves more than a simple oxidation of iron to Fe^{2+}. Iron(II) ions further undergo a reaction to form rust, which has the chemical formula $Fe_2O_3 \bullet nH_2O(s)$.

TR 1. Identify the strongest oxidizing and reducing agent in the initial reaction of iron described in the first paragraph of the passage above.

TR 2. Identify whether the change in iron that forms rust, described in the second paragraph, is an oxidation or a reduction. Support your answer.

Submit your answers to your teacher for feedback. Save a copy of your answers in your course folder, as you may wish to refer to them later in this lesson.

Read

How did you determine that iron was undergoing a change in TR 1 and TR 2? Did you focus on the charge on the iron atom? For metals, determining whether a reduction or an oxidation has occurred often requires you to look at how the charge of the metals changes. What about non-metals and atoms within polyatomic ions? How is their oxidation or reduction tracked?

Read pages 583–585 in the textbook. You may wish to save a copy of the information in "Table 1" on page 583 in your course folder.

Earlier in this module you compared the charge of copper atoms and copper ions. When doing so, you were using the oxidation numbers for these forms of the copper atom. Using oxidation numbers is a quick and useful way to obtain information about electrochemical change. Since oxidation numbers are a powerful tool in analysis of electrochemical systems and you will not always be able to access this information in your Chemistry 30 Data Booklet, you will have to memorize the information in this table.

Self-Check

SC 1. Complete "Practice" questions 1–5 on page 585 of the textbook.

Check your work with the answer in the appendix.

Read

At the beginning of this lesson you were asked to consider how the electrochemical change in atoms could be tracked. Can oxidation numbers be used to identify the oxidized and reduced species (reducing agent and oxidizing agent respectively) in a balanced chemical equation? Read the text and work through the "Sample problems" and "Communication examples" on pages 585–588 of the textbook to find out.

Self-Check

SC 2. Complete "Practice" questions 6–9 on page 588–589 of the textbook.

Check your work with the answer in the appendix.

Lesson 4 Lab: Oxidation States of Manganese

In this investigation you will use your skills of observation to investigate the changes in oxidation number that can occur with certain transitional metals. In preparation for the investigation, complete SC 3.

Self-Check

SC 3. Complete "Lab Exercise 13.B" and "Analysis" questions a–c on page 586 of the textbook.

Check your work with the answer in the appendix.

Open the Assignment Booklet to the "Module 3: Lesson 4 Assignment." You will complete "Part 1" as you view the virtual investigation "Oxidation States of Manganese" from the course multimedia DVD.

Go to the course multimedia DVD, and view the virtual investigation "Lesson 4 Lab: Oxidation States of Manganese." Remember to record data and observations as you view the presentation.

Module 3: Lesson 4 Assignment

If you have not already done so, open the Assignment Booklet. Complete "Part 1" of the "Module 3: Lesson 4 Assignment." You will receive instructions later in the lesson on when to complete the rest of the Lesson 4 Assignment.

Read

To this point you have used oxidation numbers to identify oxidized and reduced species. Read the text and work through the "Sample problems" and "Communication examples" on pages 589–593 in the textbook to see how using oxidation numbers can enable you to balance entire redox reactions.

Mastering this method will add to your skills and may save you time in completing problems in the future. You may recall that, in Lesson 3, the method to balance half-reactions and to write a net ionic equation had many steps. You may find this method preferable as an alternative or as a means to check work in which you used the other methods of writing reactions.

Self-Check

SC 4. Complete "Practice" question 12 on page 593 of the textbook.

Check your work with the answer in the appendix.

Lesson 4 Lab: Conditions Affecting the Corrosion of Iron

Problem

Do various conditions influence the corrosion of iron?

Purpose

To investigate conditions that influence the corrosion of iron objects and to determine if the chemical change observed can be explained using concepts discussed in this module.

Open the Assignment Booklet. Go to the "Module 3: Lesson 4 Assignment" and complete "Part 2: Lab—Corrosion of Nails" as you work through the following investigation. You will receive instructions later on when to complete the rest of this Assignment.

Materials

- steel wool
- acetone
- 100-mL beaker
- six nails
- six 13 mm × 100 mm test tubes
- two rubber stoppers
- test tube rack
- NaOH(aq)
- HCl(aq)
- NaCl(aq)
- deionized (distilled) water
- deaerated water
- safety glasses
- lab apron

If you have access to the materials listed, you may be able to perform this investigation.

If you do not have access to these materials and a supervised laboratory, or if you would like to see the lab performed before you attempt it, go to the course multimedia DVD, and view the virtual investigation "Lesson 4 Lab: Conditions Affecting the Corrosion of Iron." If you will not be completing the lab yourself, remember to record data and observations as you view the presentation.

Procedure

Step 1: Measure 25 mL of acetone.

Step 2: Clean six nails with steel wool, and place the nails in the acetone. Use forceps to remove the nails and then set the nails on a paper towel to dry.

Step 3: Set six 13 mm × 100 mm test tubes in a test tube rack.

Step 4: Into five of the test tubes, transfer approximately 6 mL of the following solutions: distilled water, HCl(aq), NaCl(aq), NaOH(aq), and deaerated water. Label each test tube with the name of its contents.

Step 5: Label the sixth test tube "air."

Step 6: Place a clean, dry nail in each test tube, and place a stopper in tubes 5 (dearerated water) and 6 (air).

Step 7: Set the rack containing the test tubes and nails aside for 24 hours.

Step 8: Observe the results, including solution colour, nail colour, and nail integrity. Record your observation in the data section in the "Module 3: Lesson 4 Assignment" in the Assignment Booklet.

Module 3: Lesson 4 Assignment

If you have not already done so, open your Assignment Booklet. Go to the "Module 3: Lesson 4 Assignment" and complete "Part 2." You will receive information later in the lesson on when to complete Part 3 of the Assignment.

Reflect and Connect

Billions of dollars are spent every year to prevent or slow the corrosion of iron. The following are some methods to prevent the corrosion of iron.

1. Mix metals with iron to form a more corrosion-resistant **alloy**. Alloys have a unique set of properties. Stainless steel is an iron alloy. It contains mostly iron, between 10% to 30% chromium, a smaller percentage of nickel, and sometimes cobalt.

alloy: a homogeneous mixture of two or more metals or of a metal and a metalloid

2. Paint exposed iron surfaces. Applying a coating prevents water and oxygen from gaining access to iron and therefore prevents oxidation of the iron.

3. Plate iron with another metal. Like paint, the layer of other metal prevents water and oxygen from gaining access to iron. Chromium and zinc (galvanization) are commonly used for plating iron. You will learn more about electroplating in the next module.

4. Apply oils and grease to exposed iron. This method is often used on the moving parts of machinery. Oils and grease repel water because of the lack of polarity in the molecules.

You will address a corrosion problem in the Module 3 Assessment. You might want to view the Module 3 Assessment now and consider which of the four methods listed above might be applied to the corrosion problem. You may wish to record any ideas you have so that you can revisit them after you have completed all the lessons in this module and are ready to complete the Module 3 Assessment.

You will learn more about other methods to protect iron from corrosion in Module 4.

Module 3: Lesson 4 Assignment

Open the Assignment Booklet. Complete "Part 3: Questions" of the "Module 3: Lesson 4 Assignment."

Make sure you have completed all parts of the Assignment:

- Part 1: Lab—Oxidation States of Manganese
- Part 2: Lab—Corrosion of Nails
- Part 3: Questions

Lesson Summary

In Lesson 4 you considered the following questions:

- What are oxidations numbers and how can they be used to understand redox reactions?
- What factors cause corrosion?
- How can corrosion be prevented?

Corrosion is a complex process involving more than one redox reaction. You must pay careful attention to the charge on the iron ion in order to understand the steps in the corrosion process. For other species, oxidation numbers provide insight into the chemical change of the species and the involvement of redox processes.

Lesson Glossary

alloy: a homogeneous mixture of two or more metals or of a metal and a metalloid

Module 3—Electrochemical Reactions

Lesson 5—Redox Stoichiometry

Get Focused

The label on this water treatment system says "IRON Eater."

In previous chemistry courses you learned about quantitative analysis—the determination of the quantity of substance present in a chemical system. Quantitative analysis is important in all branches of chemistry, design, and engineering.

Earlier you learned that 45 tonnes of rust were removed in the restoration of Edmonton's High Level Bridge. Now that you know more about the process of corrosion, you might ask yourself how many tonnes of iron were oxidized and, therefore, what quantity of metal needed to be replaced?

Many domestic systems also involve redox reactions. One of these is iron filtration, a commonly used water treatment method for people who rely on water from a well. In Lesson 5 you will learn about the design and use of iron filtration. You will also further your understanding of stoichiometry in a quantitative analysis of chemical systems involving redox reactions.

Consider the following questions as you complete Lesson 5:

- How can a chemical system be analyzed using redox reactions?
- How is the stoichiometric method applied to redox systems?

Module 3: Lesson 5 Assignment

You will be directed to go to the Assignment Booklet to complete Module 3: Lesson 5 Assignment later in this lesson. The assignment has the following parts:

- Part 1: Lab—Analyzing a Hydrogen Peroxide Solution
- Part 2: Reflect and Connect

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Corrosion is not the only problem iron presents. Iron is one of the most plentiful elements in Earth's crust; as such, iron can be present in groundwater. In Alberta and elsewhere, iron can be a problem in well water. Iron(II) ions give well water a pale yellow-brown colour, and iron(III) ions can create insoluble precipitates. Iron in water also gives water a metallic taste and can stain dishes and laundry.

Many people using well water install water treatment systems to remove iron from their water. A popular method to remove iron from water involves treating iron with potassium permanganate, $KMnO_4(s)$. How might this system be able to remove iron(II) ions from well water?

You may recall that permanganate ion is a strong oxidizing agent. When potassium permanganate is pumped into the filter, it brings about the oxidation of iron(II) ions to iron(III) ions. Iron(III) ions are poorly soluble in water and precipitate. The precipitate is then removed by the filter.

Self-Check

SC 1. Write the half-reactions and balanced net ionic equation for the reaction between potassium permanganate and iron(II) ions that occurs in the iron filter system. Assume acidic conditions.

SC 2. Is the reaction spontaneous?

SC 3. Potassium permanganate in an aqueous solution at low concentrations has an intense purple colour. How would a consumer know if the quantity of potassium permanganate used in the iron filter was sufficient?

Check your work with the answer in the appendix.

Read

The iron filtration system demonstrates the application of a redox reaction in a water treatment system. It also demonstrates how sensitive some processes are to the quantities of substances used.

In previous chemistry courses you studied the stoichiometric method and performed a titration experiment. Can you think of how this water treatment system is similar to a titration and a stoichiometric relationship?

Read "Titration" on page 804, and read all of page 596 of the textbook.

Self-Check

SC 4. In previous chemistry courses you performed titrations in acid-base systems. You used pH indicators to detect the endpoint of the titration.

Explain the following aspects of a titration:

a. What unique glassware is used in a titration and what is the function of each piece?

b. What is an endpoint?

c. What property of a potassium permanganate solution could be used as an endpoint for a titration?

Check your work with the answer in the appendix.

Read

You will recall from your work in previous science courses that a titration involves adding a solution with a known concentration to a test solution with an unknown concentration. When you performed a titration experiment previously, you made careful observations and you recorded the volume of each solution involved in the reaction. Using the information about the volumes in addition to the molar concentration of the titrant allows for calculation of the chemical quantities involved. You will also recall that knowledge of the balanced chemical equation for the reaction occurring in the titration is critical to performing any titration analysis.

Given the importance of writing redox reactions to performing quantitative analysis, you may understand why so much emphasis was placed on learning how to write balanced chemical equations for redox reactions earlier in this module.

To complete your work in this lesson you will need to join the following two sets of skills:

- writing redox reactions
- performing stoichiometric calculations using titration data

Carefully read and work through "Sample problem 13.14" on page 597 of the textbook to learn how to join these skills.

Self-Check

SC 5. Complete "Practice" questions 3 and 5 on page 598 of the textbook.

Check your work with the answer in the appendix.

Lesson 5 Lab: Analyzing a Hydrogen Peroxide Solution

In this investigation you will perform a titration analysis to determine the concentration of hydrogen peroxide in a commercially sold preparation.

Open the Assignment Booklet. You will record data and complete your analysis in the "Module 3: Lesson 5 Assignment."

Read page 603 in the textbook to familiarize yourself with the purpose, problem, design, materials, and procedure you will follow. Then, complete the lab as described on page 603 of the textbook.

If you do not have access to the materials and a supervised laboratory, or if you would like to see the lab performed before you attempt it, go to the course multimedia DVD, and view the virtual investigation "Lesson 5 Lab: Analyzing a Hydrogen Peroxide Solution." If you will not be completing the lab yourself, remember to record data and observations as you view the presentation.

Module 3: Lesson 5 Assignment

Go to the Assignment Booklet, and complete "Part 1" of the "Module 3: Lesson 5 Assignment." You will receive instructions later in this lesson on when to complete Part 2.

Reflect and Connect

Earlier in this lesson you learned about the use of potassium permanganate in an iron filtration system for well water.

Go to the Assignment Booklet, and complete "Part 2: Reflect and Connect" of the "Module 3: Lesson 5 Assignment."

Lesson Summary

In Lesson 5 you considered the following questions:

- How can a chemical system be analyzed using redox reactions?
- How is the stoichiometric method applied to redox systems?

An iron filtration system uses a strong oxidizing agent, permanganate ions, to oxidize iron(II) ions into a form that allows them to be filtered and removed from drinking water. The selection of permanganate as a reactant not only assures a spontaneous and stoichiometric reaction but also allows for assessment of the system's operation.

Similar principles are applied in the design of a redox titration in which spontaneous and stoichiometric reactions are used to perform quantitative analysis. Some oxidizing agents have distinctive colours, like permanganate and dichromate. As these substances undergo electrochemical change, their colour change or lack of change can provide a detectable endpoint for titrations, allowing for accurate quantitative analysis.

Module 3—Electrochemical Reactions

Module Summary

In Module 3 you considered the following module questions:

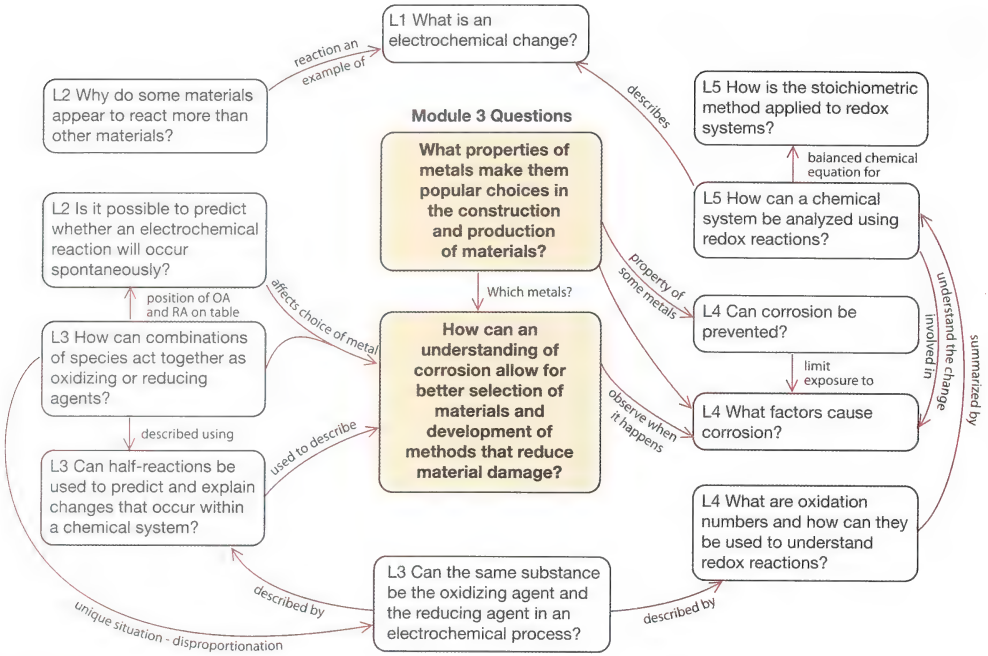
- What properties of metals make them popular choices in the construction and production of materials?
- How can an understanding of corrosion allow for better selection of materials and for development of methods that reduce material damage?

Your study of electrochemical change in Module 3 has prepared you to understand the rationale for many technologies used in products you will purchase. You also now understand the importance of maintaining or improving the methods used to protect metals from conditions in which the metals may react and fail to perform as intended.

Concept Map or Graphic Organizer

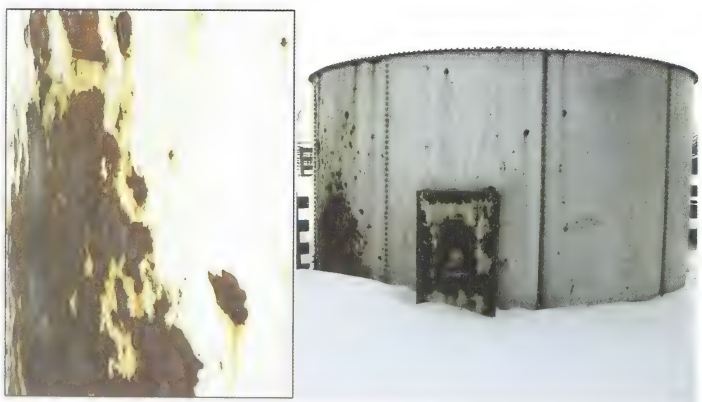
As you worked through Module 3, you may have added information to a concept map or graphic organizer based on the module and lesson questions listed in the Module 3 Concept Organizer. Now is a good time to review the relationships in your concept map or graphic organizer and to try to answer the module and lesson questions.

A sample Module 3 concept map shows one set of possible links between the questions. Remember that this is one possible description only—there are many other correct possibilities. However, if your completed concept map or graphic organizer differs significantly from the sample, you may wish to contact your teacher or to compare your map or organizer with those of other students in your class. This will ensure that your interpretations of lesson materials and your descriptions are accurate.


L1 What is an electrochemical change?
reaction an example of
L2 Why do some materials appear to react more than other materials?
Module 3 Questions
What properties of metals make them popular choices in the construction and production of materials?
describes
L5 How is the stoichiometric method applied to redox systems?
balanced chemical equation for
L5 How can a chemical system be analyzed using redox reactions?
understand the change involved in
summarized by
L2 Is it possible to predict whether an electrochemical reaction will occur spontaneously?
position of OA and RA on table
affects choice of metal
Which metals?
property of some metals
L4 Can corrosion be prevented?
limit exposure to
L3 How can combinations of species act together of oxidizing or reducing agents?
How can an understanding of corrosion allow for better selection of materials and development of methods that reduce material damage?
observe when it happens
L4 What factors cause corrosion?
described using
used to describe
L3 Can half-reactions be used to predict and explain changes that occur within a chemical system?
L4 What are oxidation numbers and how can they be used to understand redox reactions?
described by
L3 Can the same substance be the oxidizing agent and the reducing agent in an electrochemical process?
described by
unique situation - disproportionation

Module Assessment

Consider these photographs:



1. Describe the electrochemical change that has discoloured the iron used to construct the granary. Your description must mention oxidation numbers and include a relevant half-reaction.

2. List technologies that could have been used to prevent this change from occurring. Select one of the technologies you listed, and prepare a detailed description of how this method prevents this type of change.

Where possible, use appropriate terminology, concepts, reaction equations, and diagrams in your response. Save your response in your course folder and submit a copy to your teacher for assessment.

Module 3 Glossary

alloy: a homogeneous mixture of two or more metals or a metal and a metalloid

non-spontaneous: a process that is incapable of proceeding unless driven by an outside source of energy

spontaneous: a process that occurs without any external energy source

A spontaneous process will occur on its own.

Appendix

Module 3 Self-Check Answers

Contact your teacher if your answers vary significantly from the answers provided here.

Lesson 1

SC 1.

Practice 7.

a. redox reaction: a reaction in which electrons are transferred between entities

b. reduction: a chemical process involving the gain of electrons by an entity

c. oxidation: a chemical process involving the loss of electrons by an entity

Practice 8.

a. $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (gain of electrons)
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (loss of electrons)

b. $2\ H^+(aq) + 2e^- \rightarrow H_2(g)$ (gain of electrons)
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ (loss of electrons)

Practice 9.

a. $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$ (oxidation half-reaction)
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (reduction half-reaction)

b. $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$ (oxidation half-reaction)
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (reduction half-reaction)

c. $Ca(s) \rightarrow Ca^{2+}(aq) + 2e^-$ (oxidation half-reaction)
$2\ H^+(aq) + 2e^- \rightarrow H_2(g)$ (reduction half-reaction)

d. $Al(s) \rightarrow Al^{3+}(s) + 3e^-$ (oxidation half-reaction)
$Fe^{3+}(s) + 3e^- \rightarrow Fe(l)$ (reduction half-reaction)

You may have noticed that nitrate is a spectator in a, b, and c. Oxide is a spectator in d.

Practice 10.

$Cl_2(g) + 2e^- \rightarrow 2\ Cl^-(aq)$ (gain of electrons)
$2\ I^-(aq) \rightarrow I_2(s) + 2e^-$ (loss of electrons)

Practice 11.

A redox reaction has not taken place. The entities present before the reaction are Fe^{3+}, Cl^-, Na^+, and OH . Exactly the same entities are present after the reaction, so there has been no transfer of electrons from one entity to another.

SC 2.

a. As nickel is oxidized a blue-green color appears in the water because this is the color of Ni^{2+} in solution. See page 568 of the textbook for a photograph of this.

b. As copper is oxidized, two ions can be produced. Each of these ions has a different colour. Copper(I) ion has a blue-green colour, and copper(II) has a blue colour.

c. As chromium is oxidized blue is the first colour to appear. This happens as chromium is oxidized to chromium(II). If excess nitric acid exists in the system, the chromium(II) ion may be further oxidized to chromium(III) ion, which will cause the solution to become blue-green in colour.

SC 3.

oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
reduction: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$

Lesson 2

SC 1.

Practice 1.

Oxidation and reduction refer to the process of electron transfer; oxidizing and reducing agents refer to species that gain or lose electrons in an electrochemical process.

Practice 2.

Strong oxidizing agents have a strong affinity for electrons and promote the loss of electrons by other substances during a reaction.

Practice 3.

Strong reducing agents have a weak affinity for electrons and tend to lose electrons to other substances during a reaction.

Practice 5.

From "Table 1" on page 569 of the textbook or from "Table 2" on page 570 of the textbook, the metals that reacted spontaneously with copper(II) ion were lead and zinc.

Practice 6.

Silver metal did not react with copper(II) ion.

Practice 7.

In "Table 3" on page 570 of the textbook, a metal that did react spontaneously with $Cu^{2+}(aq)$ appears below $Cu^{2+}(aq)$ in the table.

Practice 8.

Question 5 using Pb^{2+} instead of Cu^{2+}: From "Table 1" or "Table 2" the metal that reacted spontaneously with lead(II) ion was zinc.

Question 6 using Pb^{2+} instead of Cu^{2+}: Silver metal and copper metal did not react with Pb^{2+}.

Question 7 using Pb^{2+} instead of Cu^{2+}: In "Table 3" the metal that did react spontaneously with Pb^{2+} appears below and to the right of Pb^{2+} in the table.

Practice 9.

Hypothesis: A spontaneous reaction will occur between a metal ion and a metal that is below it in "Table 3."

Predictions based on "Table 3" include that spontaneous reactions will occur between the following pairs:

- Ag^{+} and Cu
- Ag^{+} and Pb
- Ag^{+} and Zn
- Cu^{2+} and Pb
- Cu^{2+} and Zn
- Pb^{2+} and Zn

The predictions are correct according to the evidence collected in the experiment, so the hypothesis is supported.

Practice 10.

$Cl_2(aq) + 2e^- \leftrightarrow 2\ Cl^-(aq)$
$Br_2(aq) + 2e^- \leftrightarrow 2\ Br^-(aq)$
$I_2(aq) + 2e^- \leftrightarrow 2\ I^-(aq)$

SC 2.

Practice 11.

$Co^{2+}(aq) + 2\ e^- \leftrightarrow Co(s)$
$Zn^{2+}(aq) + 2\ e^- \leftrightarrow Zn(s)$
$Mg^{2+}(aq) + 2\ e^- \leftrightarrow Mg(s)$

Practice 12.

$Cu^{2+}(aq) + 2\ e^- \leftrightarrow Cu(s)$
$2\ H^+(aq) + 2\ e^- \leftrightarrow H_2(g)$
$Cd^{2+}(aq) + 2\ e^- \leftrightarrow Cd(s)$
$Be^{2+}(aq) + 2\ e^- \leftrightarrow Be(s)$
$Ca^{2+}(aq) + 2\ e^- \leftrightarrow Ca(s)$

Practice 13.

The spontaneity rule is based on empirical evidence from experimental results.

Practice 14.

$Cl_2(g) + 2\ e^- \leftrightarrow 2\ Cl^-(aq)$
$Br_2(g) + 2\ e^- \leftrightarrow 2\ Br^-(aq)$
$Ag^+(aq) + e^- \leftrightarrow Ag(s)$
$I_2(s) + 2\ e^- \leftrightarrow 2\ I^-(aq)$
$Cu^{2+}(aq) + 2\ e^- \leftrightarrow Cu(s)$

SC 3.

Practice 20.

a. spontaneous
b. non-spontaneous
c. non-spontaneous
d. spontaneous
e. spontaneous
f. spontaneous

Lesson 3

SC 1.

Practice 12.

a.

Step 1: $N_2O(g) \rightarrow N_2(g)$

Step 2: $N_2O(g) \rightarrow N_2(g)$ Nitrogen is balanced

Step 3: $N_2O(g) \rightarrow N_2(g) + H_2O(l)$

Step 4: $2\ H^+(aq) + N_2O(g) \rightarrow N_2(g) + H_2O(l)$

Step 5: $2^+ \quad 0 = 0 \quad {}^+0$

$2^+ = 0$

Add e^- to the most positive side.

$$\boxed{2\ e^- + 2\ H^+(aq) + N_2(g) \rightarrow N_2O(g) + H_2O(l)} \quad \text{(reduction)}$$

b.

Step 1 & 2: $NO_2^-(aq) \rightarrow NO_3^-(aq)$

Step 3: $H_2O(l) + NO_2^-(aq) \rightarrow NO_3^-(aq)$

Step 4: $H_2O(l) + NO_2^-(aq) \rightarrow NO_3^-(aq) + 2\ H^+(aq)$

$0 \quad -1 = -1 \quad +2$

$-1 = +1$

Step 5: $H_2O(l) + NO_2^-(aq) \rightarrow NO_3^-(aq) + 2\ H^+(aq) + 2\ e^-$

Step 6: $2\ OH^-(aq) + \cancel{H_2O(l)} + NO_2^-(aq) \rightarrow NO_3^-(aq) + \cancel{2\ H^+(aq) + 2\ OH^-(aq)} + 2\ e^-$

$\overset{1}{\cancel{2}}\ H_2O(l)$

$$\boxed{2\ OH^-(aq) + NO_2^-(aq) \rightarrow NO_3^-(aq) + H_2O + 2\ e^-} \quad \text{(oxidation)}$$

c.

Step 1: $Ag_2O(s) \rightarrow Ag(s)$

Step 2: $Ag_2O(s) \rightarrow 2\ Ag(s)$

Step 3: $Ag_2O(s) \rightarrow 2\ Ag(s) + H_2O_{(l)}$

Step 4: $2\ H^+(aq) + Ag_2O(s) \rightarrow 2\ Ag(s) + H_2O(l)$

Step 5: $2e^- + 2\ H^+(aq) + Ag_2O(s) \rightarrow 2\ Ag(s) + H_2O(l)$

Step 6: $2\ e^- + \cancel{2\ H^+(aq)} + \cancel{2\ OH^-(aq)} + Ag_2O(s) \rightarrow 2\ Ag(s) + \cancel{H_2O(l)} + 2\ OH^-(aq)$

$\overset{1}{\cancel{2}}\ H_2O(l)$

Step 7: $\boxed{2\ e^- + H_2O(l) + Ag_2O(s) \rightarrow 2\ Ag(s) + 2\ OH^-(aq)}$ (reduction)

d.

Step 1 & 2: $NO_3^-(g) \rightarrow HNO_2(aq)$

Step 3: $NO_3^-(g) \rightarrow HNO_2(aq) + H_2O(l)$

Step 4: $3\ H^+(aq) + NO3^-(g) \rightarrow HNO_2(aq) + H_2O(l)$

Step 5: $\boxed{2\ e^- + 3\ H^+(aq) + NO_3^-(aq) \rightarrow HNO_2(aq) + H_2O(l)}$ (reduction)

e.

Step 1 & 2: $H_2(g) \rightarrow H_2O(l)$

Step 3: $H_2O(l) + H_2(g) \rightarrow H_2O(l)$

Step 4: $H_2O(l) + H_2(g) \rightarrow H_2O(l) + 2\ H^+(aq)$

Step 5: $H_2O(l) + H_2(g) \rightarrow H_2O(l) + 2\ H^+(aq) + 2\ e^-$

Step 6 & 7: $2\ OH^-(aq) + H_2O(l) + H_2(g) \rightarrow H_2O(l) + \cancel{2\ H^+(aq) + 2\ OH^-(aq)} + 2\ e^-$

$2\ H_2O$

Simplifies: $2\ OH^-(aq) + \cancel{H_2O(l)} + H_2(g) \rightarrow \overset{2}{\cancel{3}}\ H_2O(l) + 2e^-$

$\boxed{2\ OH^-(aq) + H_2(g) \rightarrow 2\ H_2O(l) + 2\ e^-}$ (oxidation)

SC 2.

Practice 26.

a.

Species List	$Zn(s)$, $H^+(aq)$, $Cl^-(aq)$, $H_2O(l)$
Oxidizing Agents	$H^+(aq)$, $H_2O(l)$ Strongest OA = $H^+(aq)$
Reducing Agents	$Zn(s)$, $Cl^-(aq)$, $H_2O(l)$, $Cl^-(aq)$ and $H_2O(l)$ Strongest RA = $Zn(s)$
Reduction Half-Reaction	$2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$
Oxidation Half-Reaction	$Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-$
Net Equation	$2\ H^+(aq) + Zn(s) \rightarrow H_2(g) + Zn^{2+}(aq)$
The reaction is spontaneous since the position of the SOA is above the SRA in the table.	
Diagnostic Tests	For hydrogen, a "pop" test. A "pop" sound is heard when a flame is inserted into the gas produced by the reaction.
	Measure a change in pH before and after the reaction. pH is expected to increase due to the decrease in the concentration of available hydrogen ions that are used in the reaction.

b.

Species List	$Au(s)$, $H^+(aq)$, $Cl^-(aq)$, $H_2O(l)$
Oxidizing Agents	$H^+(aq)$, $H_2O(l)$ Strongest OA = $H^+(aq)$
Reducing Agents	$Au(s)$, $Cl^-(aq)$, $H_2O(l)$, $Cl^-(aq)$, and $H_2O(l)$ Strongest RA = $H_2O(l)$
Reduction Half-Reaction	$2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$
Oxidation Half-Reaction	$2\ H_2O(l) \rightarrow O_2(g) + 4\ H^+(aq) + 4\ e^-$
Net Equation	$2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$
The reaction is non-spontaneous since the position of the SOA is below the SRA in the table.	
Diagnostic Tests	No change in observable properties.

c.

Species List	$H^+(aq)$, $NO_3^-(aq)$, $Cu(s)$, $H_2O(l)$
Oxidizing Agents	$H^+(aq)$, $NO_3^-(aq)$, $H_2O(l)$ Strongest OA = $H^+(aq)$ AND $NO_3^-(aq)$
Reducing Agents	$Cu(s)$, $H_2O(l)$ Strongest RA = $Cu(s)$
Reduction Half-Reaction	$2\ NO_3^-(aq) + 4\ H^+(aq) + 2\ e^- \rightarrow N_2O_4(g) + 2\ H_2O(l)$
Oxidation Half-Reaction	$Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^-$
Net Equation	$2\ NO_3^-(aq) + 4\ H^+(aq) + Cu(s) \rightarrow N_2O_4(g) + 2\ H_2O(l) + Cu^{2+}(aq)$
The reaction is spontaneous since the position of the SOA is above the SRA in the table.	
Diagnostic Tests	Observe colour change as the reaction proceeds. A blue colour should develop as aqueous copper(II) ions are produced by the reaction. This could be confirmed by a blue colour resulting from a flame test of the solution containing the products of the reaction.

SC 3.

Practice 31.

a.

$4\ [Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-]$
$NO_3^-(aq) + 10\ H^+(aq) + 8\ e^- \rightarrow NH_4^+(aq) + 3\ H_2O(l)$

$4\ Zn(s) + NO_3^-(aq) + 10\ H^+(aq) \rightarrow 4\ Zn^{2+}(aq) + NH_4^+(aq) + 3\ H_2O(l)$

b.

$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq)$
$SO_2(g) + 2\ H_2O(l) \rightarrow SO_4^{2-}(aq) + 4\ H^+(aq) + 2\ e^-$

$Cl_2(g) + SO_2(g) + 2\ H_2O(l) \rightarrow 2\ Cl^-(aq) + SO_4^{2-}(aq) + 4\ H^+(aq)$

Lesson 4

SC 1.

Practice 1.

a. S in SO_2

$$x + 2(-2) = 0$$
$$x = +4$$

The oxidation number of S in SO_2 is +4.

b. Cl in $HClO_4$

$$+1 + x + 4(-2) = 0$$
$$x = 7$$

The oxidation number of Cl in $HClO_4$ is +7.

c. S in SO_4^{2-}

$$x + 4(-2) = -2$$
$$x = +6$$

The oxidation number of S in SO_4^{2-} is +6.

d. Cr in $Cr_2O_7^{2-}$

$$2(x)+7(-2)=-2$$
$$x=+6$$

The oxidation number of Cr in $Cr_2O_7^{2-}$ is +6.

e. I in MgI_2

The compound contains Mg^{2+} and I^-.

I^- is a monoatomic ion, so the oxidation number of I in MgI_2 is –1.

f. H in CaH_2

The compound contains Ca^{2+} and H^- and is a hydride.

The oxidation number of H in CaH_2 is –1.

Practice 2.

a. $N_2O(g)$ $2x + 1(-2) = 0$ $x = +1$

b. $NO(g)$ $x + 1(-2) = 0$ $x = +2$

c. $NO_2(g)$ $x + 2(-2) = 0$ $x = +4$

d. $NH_3(g)$ $x + 3(+1) = 0$ $x = -3$

e. $N_2H_4(g)$ $2x + 4(+1) = 0$ $x = -2$

f. $NaNO_2(s)$ contains Na^+ and NO_3^-, using NO_3^-, $x + 3(-2) = -1$ $x = +5$

g. $N_2(g)$ is an element. The oxidation number is 0.

h. $NH_4Cl(s)$ contains NH_4^+ and Cl^-, using NH_4^+, $x + 4(+1) = +1$ $x = -3$

Practice 3.

a. C(s) (graphite) is an element, so the oxidation number is 0.

b. $C_6H_{12}O_6$ $6x + 12(+1) + 6(-2) = 0$ $x = 0$

c. Na_2CO_3 contains Na^+ and CO_3^{2-}.
using CO_3^{2-}, $x + 3(-2) = -2$ $x = +4$

d. CO $x + 1(-2) = 0$ $x = +2$

Practice 4.

Remember that coefficients in an equation do not affect the value of oxidation numbers.

O in ClO_3^-, ClO_2, and H_2O is –2.

H in H^+ and H_2O is +1.

Cl in ClO_3^- is +5 ($x + 3(-2) = -1$).

Cl in Cl^- is –1.

Cl in ClO_2 is +4 ($x + 2(-2) = 0$).

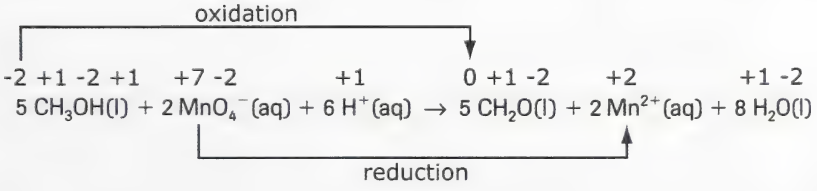
Cl in Cl_2 is 0 (element).

Practice 5.

CH_4	$x + 4(+1) = 0$	$x = -4$
CH_3OH	$x + 1(-2) + 4(+4) = 0$	$x = -2$
CH_2O	$x + 2(+1) + 1(-2) = 0$	$x = 0$
HCOOH	$x + 2(+1) + 2(-2) = 0$	$x = +2$
CO_2	$x + 2(-2) = 0$	$x = +4$

SC 2.

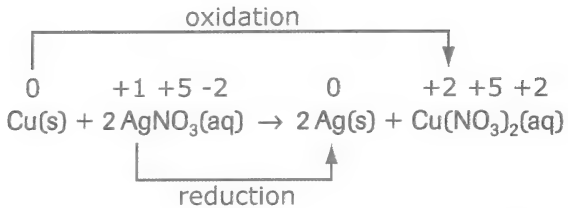
Practice 6.



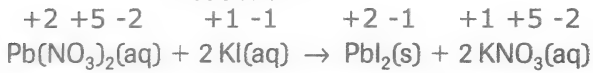
b. Carbon is oxidized from –2 in methanol to 0 in methanal.

c. Manganese is reduced from +7 in permanganate ion to +2 in manganese(II) ion.

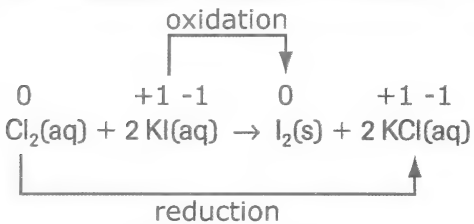
Practice 7.

a. 

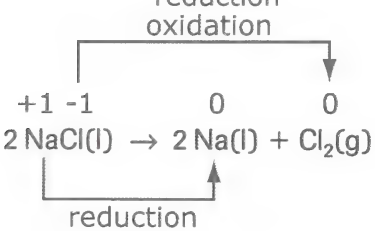
(redox)

b. 

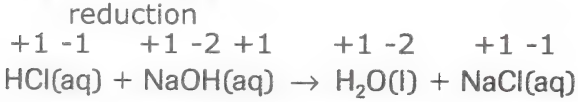
(not redox)

c. 

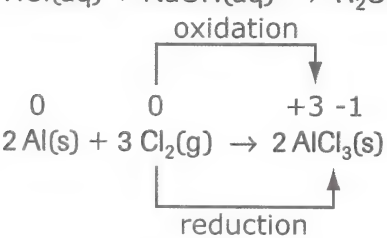
(redox)

d. 

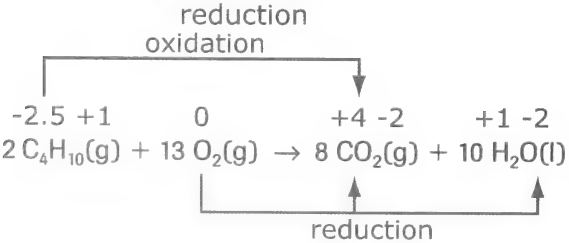
(redox)

e. 

(not redox)

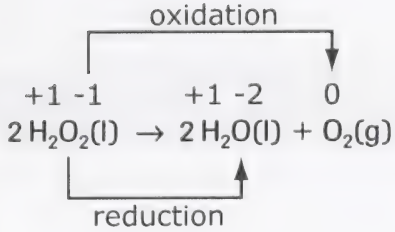
f. 

(redox)

g. 

(redox)

h.



(redox)

Practice 8.

a. single replacement
b. double replacement
c. single replacement
d. simple decomposition
e. double replacement
f. formation
g. complete combustion
h. decomposition (but not simple)

Double replacement reactions do not appear to be redox reactions.

Practice 9.

In $H_2O_2(l)$, the oxidation number of oxygen is –1, because the compound is a peroxide (see "Table 1" on page 583 of the textbook). If it reacts with a strong-enough oxidizing agent, it will be converted to $O_2(g)$ and $H^+(aq)$, and the oxidation number of oxygen will increase from –1 to 0, which is an oxidation process. If $H_2O_2(l)$ reacts with a strong-enough reducing agent, it will be converted to $H_2O(l)$, and the oxidation number of oxygen will decrease from –1 to –2, which is a reduction process.

SC 3.

a.

Reaction in Table 3	Vanadium Ions in Sequence of Reactions
1	$VO_3^-(aq)$
2	$VO_3^-(aq) \rightarrow VO^{2+}(aq) \rightarrow V^{3+}(aq) \rightarrow V^{2+}(aq)$
3	$V^{2+}(aq) \rightarrow V^{3+}(aq)$
4	$VO_3^-(aq) \rightarrow VO_2^+(aq)$
5	$VO_2^+(aq)$ (no change)
6	$V^{2+}(aq) \rightarrow V^{3+}(aq) \rightarrow VO_2^+(aq) \rightarrow VO_3^-(aq)$

b.

Reaction in Table 3	Description of Chemical Change
1	Ammonium vanadate(V) dissociates, thus no chemical change or change in oxidation number occurs.
2	Vanadium is reduced from +5 to +2.
3	Vanadium is oxidized from +2 to +3.
4	Vanadium is reduced from +5 to +4.
5	There is no change in oxidation number.
6	Vanadium is oxidized from +2 to +5.

c.

Reaction in Table 3	Explanation of Chemical Change
3	Vanadium is oxidized from +2 to +3, probably due to exposure to atmospheric oxygen.
4	Vanadium is reduced from +5 to +4, forming $VO_2^+(aq)$. The blue colour of the $VO_2^+(aq)$ coincides with the appearance of aqueous iodine, the product of the other half-reaction in the system. Aqueous iodine is brown in colour. The combination of the two colours results in the solution appearing black.
5	There is no change in oxidation number because iodide ions do not react spontaneously with $VO_2^+(aq)$.
6	Vanadium is oxidized from +2 to +5 forming $VO_3^-(aq)$, since permanganate ions in acidic conditions react spontaneously with $V^{2+}(aq)$.

SC 4.

Practice 12.

a.

$$14\ H^+(aq) + Cr_2O_7^{2-}(aq) + 6\ Cl^-(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ Cl_2(aq) + 7\ H_2O(l)$$

$3e^-/Cr$ $\quad$ $1e^-/Cl^-$

$6e^-/Cr_2O_7^{2-}$ $\quad$ $1e^-/Cl^-$

$\times 1$ $\quad$ $\times 6$

$+5\ -2 \quad +1\ +4\ -2 \quad +6\ -2 \quad 0 \quad +1\ -2 \quad +1$

b.

$$2\ IO_3^-(aq) + 5\ HSO_3^-(aq) \rightarrow 5\ SO_4^{2-}(aq) + I_2(s) + H_2O(l) + 3\ H^+(aq)$$

$5e^-/I$ $\quad$ $2e^-/S$

$5e^-/IO_3^-$ $\quad$ $2e^-/HSO_3^-$

$\times 2$ $\quad$ $\times 5$

$+1\ -1 \quad +1\ +6\ -2 \quad +4\ -2 \quad 0 \quad +1\ -2$

c.

$$2\ HBr(aq) + H_2SO_4(aq) \rightarrow SO_2(g) + Br_2(l) + 2\ H_2O(l)$$

$1e^-/Br$ $\quad$ $2e^-/S$

$1e^-/HBr$ $\quad$ $2e^-/H_2SO_4$

$\times 2$ $\quad$ $\times 1$

After balancing a net ionic equation using oxidation numbers, you should check to see if the charge on each side balances.

Lesson 5

SC 1.

$$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$

$$5\ [Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + 1\ e^-]$$

$$5\ Fe^{2+}(aq) + MnO_4^-(aq) + 8\ H^+(aq) \rightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(l)$$

SC 2. The reaction is spontaneous.

SC 3. When permanganate reacts, it undergoes reduction to form $Mn^{2+}(aq)$. The manganese(II) ion at low concentrations is colourless. If appropriate quantities of permanganate are being used, the water from the iron filter will appear colourless. If excessive permanganate is used, the water will appear pink or purple in colour. If not enough permanganate is used, the water will appear pale yellow-orange in colour due to the presence of iron(II) ions.

SC 4.

a. A burette (to measure volume of titrant added to reach endpoint) and a pipette (to deliver a specific volume of solution of unknown concentration (test solution)) are used in a titration.

b. The endpoint is a point in a titration when an observable change in the system occurs. Usually this change is a colour change, but it can be another change such as a change in pH or conductivity.

c. Colour could be used as an endpoint. The persistence of the purple $MnO_4^-(aq)$ indicates an excess of permanganate ion in the solution. Permanganate ion can only exist in excess in the solution once all of the reactant in the test solution has reacted.

SC 5.

Practice 3.

$Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$ (oxidation)

$(Ag^+(aq) + 1e^- \rightarrow Ag(s)) \times 2$ (reduction)

$Ni(s) + 2\,Ag^+(aq) \rightarrow Ni^{2+}(aq) + 2\,Ag(s)$ (redox)

$$\begin{aligned}\text{mol Ni} &= \frac{m}{M} \\ &= \frac{25.0\text{ g}}{58.69\text{ g/mol}} \\ &= 0.426\text{ mol}\end{aligned}$$

$$\begin{aligned}\text{mol Ag}^+ &= 0.426\text{ mol} \times \frac{2}{1} \\ &= 0.852\text{ mol}\end{aligned}$$

$$\begin{aligned}\text{vol Ag}^+ &= \frac{n}{c} \\ &= \frac{0.852\text{ mol}}{0.10\text{ mol/L}} \\ &= 8.5\text{ L}\end{aligned}$$

The volume of 0.10 mol/L silver ion solution that will react is 8.5 L.

Practice 5.

First, calculate the concentration of the Fe^{2+} primary standard solution:

$$\begin{aligned}\text{mol } Fe^{2+} &= \text{mol Fe} \\ &= \frac{m}{M} \\ &= \frac{1.08\text{ g}}{55.85\text{ g/mol}} \\ &= 0.0193\text{ mol}\end{aligned}$$

$$\begin{aligned}\left[Fe^{2+}\right] &= \frac{n}{V} \\ &= \frac{0.0193\text{ mol}}{0.2500\text{ L}} \\ &= 0.0774\text{ mol/L}\end{aligned}$$

Titration

$(Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + 1e^-) \times 5$ (oxidation)

$MnO_4^-(aq) + 8\,H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4\,H_2O(l)$ (reduction)

$5\,Fe^{2+}(aq) + MnO_4^-(aq) + 8\,H^+(aq) \rightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l)$ (redox)

$$\begin{aligned}\text{mol } Fe^{2+} &= cV \\ &= (0.0774\text{ mol/L})(0.0100\text{ L}) \\ &= 0.000\,774\text{ mol}\end{aligned}$$

$$\begin{aligned}\text{mol } MnO_4 &= 0.000\,774\text{ mol} \times \frac{1}{5} \\ &= 0.000\,155\text{ mol}\end{aligned}$$

$$\begin{aligned}\left[MnO_4^-\right] &= \frac{n}{V} \\ &= \frac{0.000\,155\text{ mol}}{0.0136\text{ L}} \\ &= 0.0114\text{ mol/L}\end{aligned}$$

The concentration of the permanganate solution is 0.0114 mol/L.

Chemistry 30

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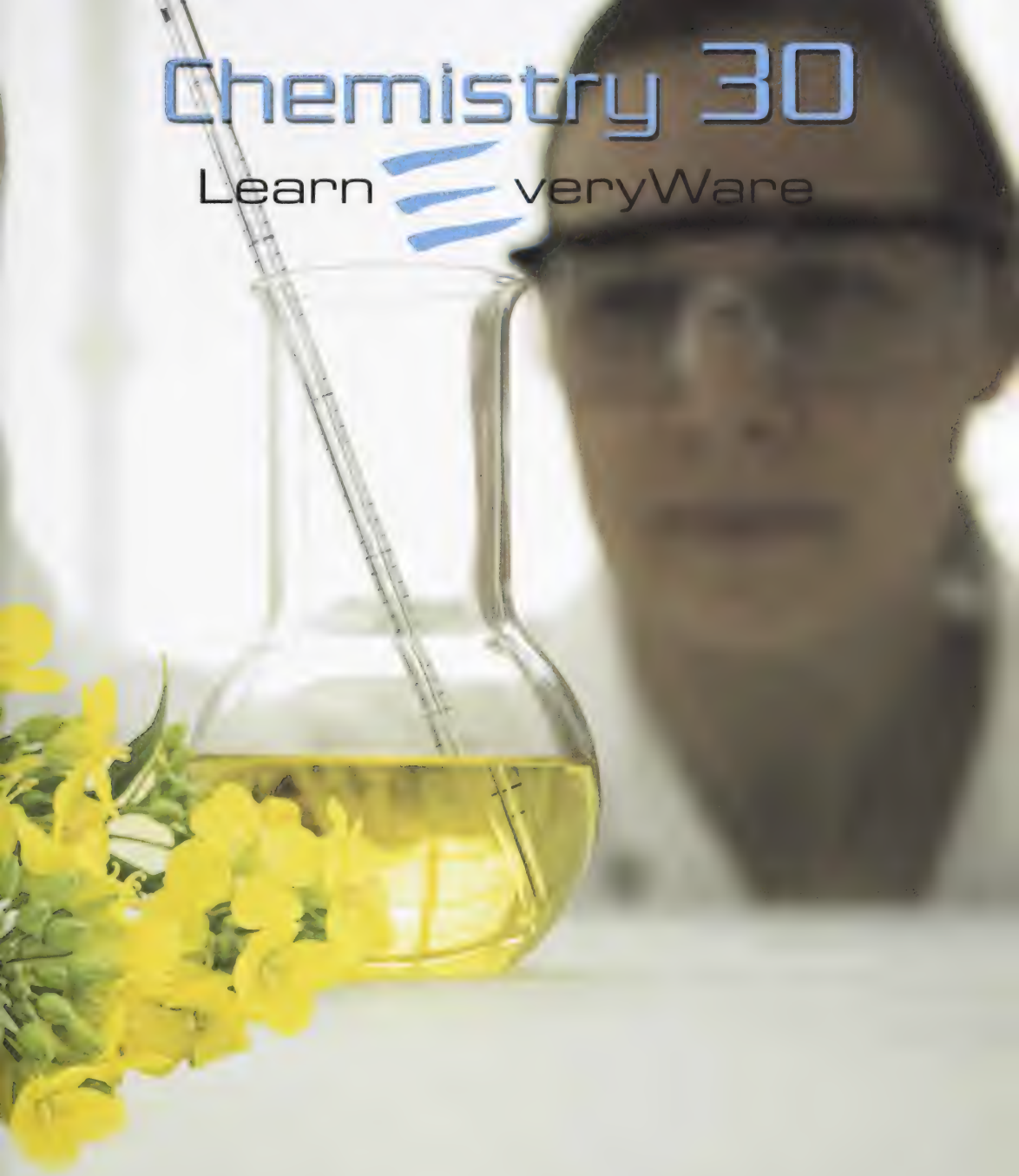

Batteries and Balance

Module 4

Contents

Module 4 Introduction 2

Big Picture 3

In This Module 4

Lesson 1: Voltaic Cells and Batteries 6

Lesson 2: Design of Commercial Cells 13

Lesson 3: Electrolytic Cells 20

Lesson 4: Commercial Applications of Electrolytic Cells 26

Lesson 5: Quantitative Relationships in Cells 30

Module Summary 34

Module Assessment 36

Unit B Conclusion 38

Appendix (Self-Check Answers) 42

Module 4—Batteries and Balance

Module Introduction

When you look at this picture, do you see many different kinds of batteries or do you see many different shapes of the same chemical system? Your answer might indicate whether you view batteries from a technological perspective or from a scientific perspective.

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In this module you will extend and apply your understanding of electrochemical change to understanding the design and function of a variety of types of electrochemical cells, one of which is the electric cell—the technology used in batteries and in other applications.

You will use the skills and techniques you have been developing to analyze and predict change in electrochemical systems. You will do this within the context of electric and electrolytic cells. You will also investigate the technological aspects of, and your reliance on, these chemical systems.

As you investigate the electrochemical cells introduced in this module, you will see how chemical knowledge about reduction-oxidation reactions is applied to solving problems and expanding capabilities. You will consider the purpose and impact of different types of electrochemical cells.

In Module 4 you will investigate the following question:

- How are principles of oxidation and reduction applied to electrochemical cells?

Remember that each lesson will also be organized around questions intended to guide your study. As you proceed through Module 4, you may record answers to these questions and any interrelationships that exist between them in a concept map or graphic organizer. More information is available in the Unit B Concept Organizer. In the Module 4 Summary you will receive further information on how you can use your concept map or graphic organizer to review the concepts you studied in this module.

Big Picture

It's the end of a long day. Your end-of-day routine probably includes checking your cell phone, your MP3 player, and a few other devices to see if their batteries need to be recharged. You probably haven't considered all the other devices in your home that rely on batteries in order to work.

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The batteries in a laptop computer or in a cordless telephone provide you with great convenience, but at what cost? While some batteries are rechargeable, others are not. Every battery has a certain lifespan. When a battery reaches the end of its lifespan, it is usually just thrown away without further thought. However, batteries that end up in landfills can leak toxic chemicals into the environment.

If you ever make a trip to the Eco Station or special waste handling facility in your area, you might be amazed by how many electronic devices are disposed of. What about the batteries that powered these devices? Is it time to think about your use of batteries and to make some changes in how you use them?

In Module 4 you will learn about the design and operation of electrochemical cells. You will apply your understanding of electrochemical change learned in Module 3. You will continue to use tools like the "Table of Selected Standard Electrode Potentials," your ability to predict spontaneous and non-spontaneous reduction-oxidation reactions, and stoichiometry throughout Module 4.

Assessment in This Module

Each lesson contains a range of activities and assessment options. These include assignments, labs, and Self-Check, Try This, Discuss, Reflect and Connect, and Reflect on the Big Picture activities. Instructions will be provided for each of these activities so that you can appropriately

focus your time and effort. Your teacher will tell you which assessment options to complete and which responses to submit for marks or feedback. Remember to save all of your work in your Chemistry 30 folder.

In the Module 4 Assessment you will apply your knowledge of electrochemical systems to investigate a cell constructed using a potato or a lemon.

You may wish to look at the Module Assessment and the Unit Assessment before starting Lesson 1.

In This Module

Lesson 1—Voltaic Cells and Batteries

In Lesson 1 you will learn about the scientific principles involved in the construction and function of voltaic cells, a type of commercial electric cell.

You will investigate the following lesson questions:

- How much do you rely on batteries in your everyday life?
- What are the components of a voltaic cell?

Lesson 2—Design of Commercial Cells

In Lesson 2 you will investigate the construction of different types of commercial and consumer cells used by society, and you will consider the environmental impact of your personal battery use.

You will investigate the following lesson question:

- Why are there so many different designs and types of electrochemical cells?

Lesson 3—Electrolytic Cells

Lessons 1 and 2 will explain the chemical change that occurs within a voltaic cell. But how can you make a cell rechargeable? In Lesson 3 you will learn about rechargeable cells and why some cells are rechargeable and others are not.

You will investigate the following lesson questions:

- What is an electrolytic cell?
- What chemical and energy conversions occur within an electrolytic cell?

Lesson 4—Commercial Applications of Electrolytic Cells

In Lesson 4 you will learn about industrial applications of electrolysis, and you will learn how this technology can be applied to the extraction of elements and to the refining and electroplating of metals.

You will investigate the following lesson question:

- What are some of the practical applications of electrolytic cells?

Lesson 5—Quantitative Relationships in Cells

In Lesson 5 you will record and analyze qualitative changes within electric and electrolytic cells.

You will investigate the following lesson question:

- What are the important quantitative relationships within operating electric and electrolytic cells?

Module 4—Batteries and Balance

Lesson 1—Voltaic Cells and Batteries

Get Focused

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The commercial "batteries" you use are made up of components involved in an electrochemical reaction. In Module 3 you learned about the scientific principles associated with electrochemical change. In this lesson you will begin to investigate how materials can be used to make model systems and commercial cells that allow for the study and use of the electron exchange that occurs between components.

You will start your investigation of electrochemical cells by looking at the most-used type of electric cell—the voltaic cell. By learning more about the voltaic cell, you might even be able to identify ways to improve this technology and its application.

Consider the following questions as you complete Lesson 1:

- How much do you rely on batteries in your everyday life?
- What are the components of a voltaic cell?

Module 4: Lesson 1 Assignment

There is no assignment for this lesson.

There are other questions in this lesson that are not marked by the teacher; however, you should still answer these questions. The Self-Check, Try This, and other types of questions are placed in this lesson to help you review important information and build key concepts that may be applied in future lessons. You should record the answers to all the questions in the lesson and place those answers in your course folder.

After a discussion with your teacher, you must decide what to do with the questions that are not part of your assignment. For example, you may decide to submit the responses to Try This and

other questions that are not marked to your teacher for informal assessment and feedback. Your answers are very important to your teacher. They provide your teacher with information about your learning, and they help your teacher identify where adjustments to your instruction may be necessary.

Explore

Try This

Battery Audit

Perform an audit of the commercial electric cells, or "batteries," used in your home. You might identify all of the devices powered by batteries that you or your family recently used, or you might count the number and types of cells you replaced recently.

TR 1. In a table or spreadsheet, list each type of cell and its quantity. You might also want to indicate what device the cell came from (e.g., alarm clock) and, if you can recall, the last time the cell was changed in the device. For devices that you place on a recharger (e.g., cell phone), you might want to indicate the components of the cell (e.g., lithium ion, nickel-cadmium, metal hydride) and how often you recharge the cell. Finally, include an estimate of the quantity of cells you would use in one year.

Save your work in your course folder.

Discuss

Place a copy of your audit in the discussion area for your class. Look at the totals submitted by other students. How would you rank your household's use of batteries relative to that of other households?

Try This

In previous chemistry courses you became aware of chemical waste, which is the product of chemical processes. In this module you will learn more about the contents of commercial cells and the products of the chemical reactions that occur within them. Now is a good time to think about the mass of chemical waste products that come from commercial cells.

TR 2. Prepare a spreadsheet to calculate the mass of each type of commercial electric cell that is disposed of by two of the following groups of people. You might use the totals in your audit or in the audits of your classmates to determine average use where required.

a. your household
b. your class
c. your local community

Cell Type	Mass (g)
AAA	11.2
AA	23.9
C	70.0
9-volt	46.6
D	149.9
coin type	3.1

TR 3. How does your household dispose of spent cells? Survey other students in your class to see if there are other ways to dispose of spent cells.

Save a copy of your answers in your course folder. You may be asked to submit a copy to your teacher for feedback.

Read

Electrochemical cells are devices built to use the transfer of electrons that drives redox reactions. The components of an electrochemical cell are the reactants for a reduction-oxidation process. The design allows the exchange of electrons between reactants to do work. Keep this principle in mind as you read pages 612–614 in the textbook.

Module 4: Lesson 1 Lab—The Voltaic Cell

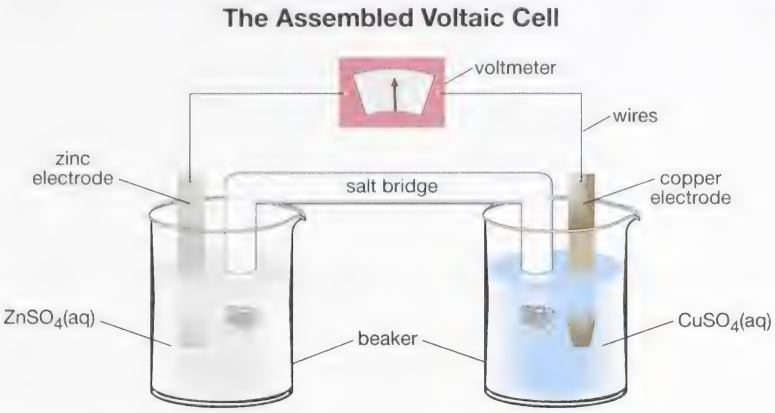
Alessandro Volta (1745–1827) is credited with the discovery of the construction and design of the first electric cells. The name *voltaic* is often used to describe the type of cell you will study in this module. One type of voltaic cell design uses copper and zinc electrodes. This type of cell is often studied due to its reliability. In this investigation you will construct and test a voltaic cell.

Pre-Lab

Earlier in this lesson you learned that the principles of electrochemical reactions you investigated in Module 3 could be applied to understand the operation of the cells you will investigate in Module 4. Before you view the virtual investigation for this lab, complete Self-Check questions 1–7.

Self-Check

Consider this diagram:



SC 1. List the two metals used in the cell. Use the "Table of Selected Standard Electrode Potentials" in your Chemistry Data Booklet to determine which one of the two metals is the more reactive metal. Repeat this step for the metal ions in the cell.

SC 2. Use the "Table of Selected Standard Electrode Potentials" in your Chemistry Data Booklet to determine the oxidation and reduction half-reactions that would occur in the voltaic cell illustrated.

SC 3. Use your knowledge of spontaneous reactions to determine whether the more reactive metal and metal ion will react spontaneously in the cell shown in the diagram.

SC 4. Use your half-reactions to determine which metal is losing electrons and which metal ion is gaining electrons.

SC 5. Use your spontaneous half-reactions to predict what will happen to each metal strip if you were to let the voltaic cell sit for a long time.

SC 6. If the voltaic cell kept operating indefinitely, would any part of the cell labelled in the diagram need to be replaced? Explain your reasoning.

SC 7. Hypothesize which metal strip is the negative electrode of the voltaic cell and which metal strip is the positive electrode of the voltaic cell. Provide a reason for your choice.

Check your work with the answer in the appendix.

Procedure

Go to the course multimedia DVD, and view the virtual investigation "Building a Voltaic Cell." Remember to record your data and observations as you view the presentation.

Analysis

Use your observations from the virtual investigation to complete Self-Check questions 8–11.

Self-Check

SC 8. Describe the effect on the output of the voltmeter when you lifted the salt bridge out of the cell.

SC 9. Hypothesize why the salt bridge is necessary for the voltaic cell to work.

SC 10. Explain what happened to the output of the voltmeter when you connected the leads from the voltmeter to the opposite electrodes in the cell.

SC 11. Describe the changes you observed to the electrodes or to the solutions. Use the half-reactions developed in SC 2 to help suggest reasons for the changes you observed.

Check your work with the answer in the appendix.

Read

In the investigation you just completed you identified the parts of a voltaic cell and, in combination with your pre-lab work, you identified reactions that involve different components of the cell. Read page 622 and the top of page 623 of the textbook to learn more about other necessary components of voltaic cells and about how voltaic cells are represented.

Try This

TR 4. A silver-copper voltaic cell is constructed.

a. Draw a diagram of the cell. Label the location of the following components: Ag(s), $AgNO_3(aq)$, Cu(s), $Cu(NO_3)_2(aq)$, salt bridge, electrical wire.

b. Write the reduction half-reaction occurring in this cell.

c. Write the oxidation half-reaction occurring in the cell.

d. Label the anode and cathode of the cell.

e. Use an arrow to indicate the direction of the flow of electrons along the electrical wire.

TR 5. The salt bridge, or porous boundary, is an essential part of a voltaic cell. Since electrons are exchanged between the half-cells, ions must also be exchanged in order to conserve charge.

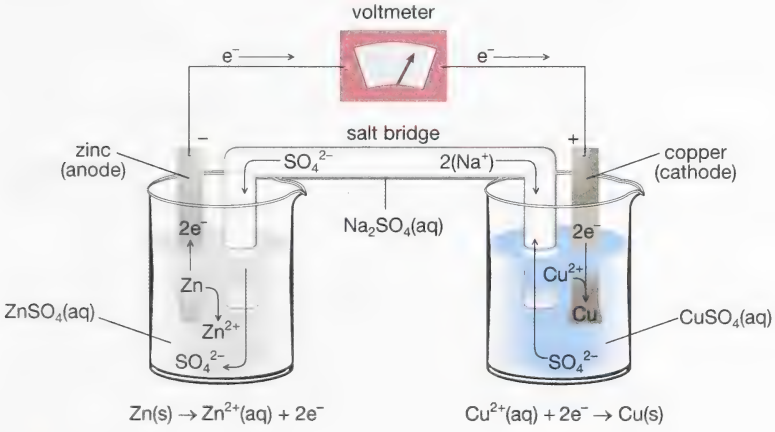
a. Predict the direction of the flow for positively charged ions (cations) with respect to the cathode half-cell.

b. Predict the direction of the flow for negatively charged ions (anions) with respect to the anode half-cell.

Watch and Listen

The Voltaic Cell

Got to the course multimedia DVD, and view the animation "The Voltaic Cell." Use the information in the animation to check your answers to questions TR 4 and TR 5. You may also wish to read pages 623–624 in the textbook for another description of the operation of, and changes occurring within, a silver-copper voltaic cell.



Read

Voltaic cells can also be constructed with half-cells that do not involve a metal as a reactant. An inert electrode can be used if a solution contains the strongest oxidizing agent or the strongest reducing agent in an aqueous form. The term *inert* means that the electrode cannot react during the cell's operation. An electrode is only necessary to be an electrical conductor—to allow for the movement of electrons in a half-cell.

Read pages 625–626 in the textbook to learn how an inert electrode is used in a voltaic cell. Work through "Communication example 1" on page 625 to check your ability to represent the use of an inert electrode in a voltaic cell.

Self-Check

SC 12. Complete "Practice" questions 1–9 on page 626 of the textbook.

Check your work with the answer in the appendix.

Reflect and Connect

In this lesson you learned to construct a voltaic cell. You also learned about the scientific principles involved in the design and operation of these cells. Read "Did You Know?" on page 615 and "Consumer, Commercial, and Industrial Cells" on pages 615–618 of the textbook. Do the different kinds of cells described in these readings have similar components to the voltaic cells you investigated in this lesson?

RC 1. Prepare a chart summarizing the similarities and differences you identified.

Save your chart in your course folder.

Lesson Summary

In Lesson 1 you considered the following questions:

- How much do you rely on batteries in your everyday life?
- What are the components of a voltaic cell?

At the beginning of this lesson you conducted an audit of the different types of commercial cells used in your house. Did your reliance on commercial cells surprise you or did you just take for granted that battery use is part of modern-day life? How were you impacted when you calculated the mass of chemical waste generated by spent commercial cells disposed of by your class or by your community?

In this lesson you learned about the parts of a voltaic cell and how these parts work together to convert chemical energy into electrical energy. You learned how to use the "Table of Selected Standard Electrode Potentials" (also used in Module 3) to predict the strongest oxidizing and reducing agents in the system. You used the information from this table to write the half-reactions that occur while the cell is functioning. You also completed a virtual investigation in which you explored the connections between half-cells and the observable changes that occur when a voltaic cell operates.

Module 4—Batteries and Balance

Lesson 2—Design of Commercial Cells

Get Focused

Wow! There are a lot of different kinds of commercial electric cells available. Some cells, like a 9-volt cell, are unique because of their high voltage. But most of the cells you purchase have a voltage of 1.5 V. In addition to voltage, there are cells made with alkaline, rechargable, lithium ion, and metal hydride components.

In this lesson, you will investigate the construction of different types of commercial electric cells and other kinds of cells used commercially. As you proceed through this lesson, you will further consider the environmental impact of your personal use of commercial electric cells. You will collect information about how substances from discarded cells might enter the environment, and you will also collect information about mechanisms that can be used to reduce the quantity of material from these cells that goes into landfills.

Consider the following question as you complete Lesson 2:

- How do you determine the cell potential for a voltaic cell?

Module 4: Lesson 2 Assignment

In the Lesson 2 Assignment you will complete the investigation "Building Voltaic Cells." You will be directed to go to the Assignment Booklet to complete the Module 4: Lesson 2 Assignment later in this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

When you look at different commercial electric cells you may notice that the cells come in many different shapes. While this can mean they have different voltages, in many cases (including in the picture shown) the cells have the same voltage. How is it possible to have cells with different shapes and the same voltage?

In Lesson 1 you constructed a zinc-copper voltaic cell, and you observed that it had a voltage of 1.10 V. You also analyzed the voltage of a copper-silver voltaic cell. Would you expect the copper-silver cell to have the same voltage as the zinc-copper cell?

As you may suspect, the choice of chemicals in your cell has a large impact on the cell's voltage. Do you think there might be a relationship between the voltage produced by a cell and the position of the reactants in the table of half-reactions you have been using throughout Unit B?

Read pages 627–630 in the textbook to learn more about voltages—standard cell potentials—produced by cells and the standard reduction potentials listed on the "Table of Selected Standard Electrode Potentials." Then work through "Communication example 2" and "Communication example 3" on page 631 of the textbook.

Self-Check

SC 1. Calculate the standard cell potential of a zinc-copper voltaic cell like the one you analyzed in Lesson 1.

SC 2. Calculate the standard cell potential of a copper-silver voltaic cell.

SC 3. Comment on the following observation: "A higher cell potential is observed with reactants that are greatly separated on the table of half-reactions."

SC 4. Suggest a reason why the three types of consumer cells shown in the picture above have the same voltage, or cell potential.

Check your work with the answer in the appendix.

Read

Sometimes the measured cell potential is not the same as the predicted standard cell potential. Read page 632 in the textbook to discover reasons for discrepancies you may encounter.

Self-Check

SC 5. Complete "Practice" questions 12–16 on page 633 of the textbook.

Check your work with the answer in the appendix.

Read

Chances are that the most-used type of commercial electric cell you found in the "battery" audit you conducted in Lesson 1 was a type of dry cell. In the Reflect and Connect section of Lesson 1 you read about "dry cells," cells in which the quantity of water within the cell has been reduced, but the properties of the electrolytes have been maintained.

A modification of the dry cell is the alkaline dry cell. The modification did not affect the measured cell potential, but it did improve the performance of the cell. Go to "Table 2" on page 616 of the textbook and read the information comparing the dry cell to the alkaline dry cell. Can you classify the "Characteristics and uses" for the alkaline dry cell listed in the table according to perspective?

Alkaline cells use KOH as an electrolyte, which forms a basic, or alkaline, solution. Recall that as the concentration of OH^- increases in a solution, the concentration of H^+ is reduced. Basic electrolytes, therefore, reduce the possibility that H^+ ions will act as an oxidizing agent in the battery. In other words, alkaline batteries tend to last longer before "going dead."

Alkaline cells typically produce a potential that is comparable to a zinc-carbon cell (approximately 1.5 V). However, the electrolyte in a zinc-carbon cell is slightly acidic. This makes the zinc-carbon cell somewhat inefficient because both H^+ ions and the intended oxidizing agent are reduced. Because of the low concentration of H^+ ions and the undesirable reaction they promote, alkaline cells can deliver significantly higher currents. This combination of higher current and longer life is a very desirable characteristic for batteries that run devices like CD players, calculators, and smoke alarms.

Consumer cells are most often evaluated in terms of their durability and convenience. When it comes to batteries, social perspectives, which demand that portable communication devices like cellular telephones stay connected, have a much higher priority than, say, environmental perspectives, which consider the consequences of using all these batteries.

Reflect on the Big Picture

Open the Assignment Booklet. You will record your response to Reflect on the Big Picture in "Part 1" of "Module 4: Lesson 2 Assignment."

Look at the picture shown. Have you ever seen a cell in this condition? Sometimes an alkaline dry cell may leak contents as a result of a faulty seal in its construction. How should you deal with a leaking cell of this type? Use a variety of resources (e.g., textbook, Materials Safety Data Sheets, Internet) to research how a leaking alkaline dry cell should be safely removed from an electronic device and disposed of. Prepare disposal instructions that you could provide to a family member.

Save any information about concerns or procedures for disposal of waste from alkaline dry cells or other cells in your course folder. You may refer to this information later in this lesson as you consider the environmental impact of using these cells.

Module 4: Lesson 2 Assignment

If you have not already done so, open the Assignment Booklet, and complete "Part 1" of "Module 4: Lesson 2 Assignment" now.

Module 4: Lesson 2 Lab—Building Voltaic Cells

In this module you have learned a great deal about the construction of voltaic and other types of electrochemical cells. In this investigation you will construct voltaic cells, and you will use the techniques you have learned to test predictions that can be made about voltaic cells.

Open the Assignment Booklet. Complete the "Part 2" Pre-Lab question in "Module 4: Lesson 2 Assignment" now. You will complete the remainder of "Part 2" after viewing the virtual investigation.

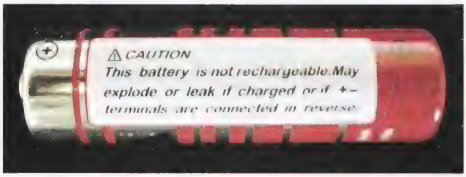
Go to the course multimedia DVD, and view the virtual investigation "Building Voltaic Cells." Remember to record your data and observations as you view the presentation.

Module 4: Lesson 2 Assignment

If you have not already done so, open the Assignment Booklet. Complete all of "Part 2" in the section "Module 4: Lesson 2 Assignment."

Read

Common Types of Commercial Cells



The warning on this battery reads "Caution. This battery is not rechargeable. May explode or leak if charged or if +- terminals are connected in reverse."

Rechargeable cells are also popular in many applications including cellular telephones, video cameras, and personal music players.

One type of rechargeable cell is the nickel-cadmium cell. The advantage of nickel-cadmium batteries is that they have a relatively low internal resistance and can therefore supply high surges of current when necessary. They are also relatively easy to recharge. The main disadvantage of nickel-cadmium cells is that they typically have a lower cell potential—1.2 V as compared to alkaline dry cells, which have a measured potential of 1.5 V.

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Alkaline cells and other cells using zinc and manganese oxide as reactants are not rechargeable; some cells even come with warnings to prevent users from trying to recharge them. Later in this module you will learn more about the process of recharging cells and about why this process cannot be accomplished with all consumer cells.

Lithium ion cells are popular in small, portable devices like cellular telephones and digital cameras, which require short bursts of significant amounts of electrical energy. Lithium ion cells are typically composed of a lithium electrode, a graphite electrode (a form of carbon), and one of several possible electrolytes.

The most common wet cells in use today are lead-acid batteries used in automobiles. Lead-acid batteries are useful for automobiles because they are relatively cheap and easily recharged. The disadvantage of a lead-acid battery relates to its chemical contents—lead and concentrated sulfuric acid. Both substances can cause harm. Considerable caution needs to be exercised when this type of battery is transported, handled, and disposed of.

Reflect on the Big Picture

In Lesson 1 you completed an audit of the cells used in your home. In this lesson you investigated common types of consumer cells.

RBP 1. Retrieve your "battery" audit and classify the cells you counted into the following categories: rechargeable and non-rechargeable.

Many cells and batteries are discarded improperly. In this lesson you learned that some of the components of various types of cells can be toxic or may cause harm to the environment.

RBP 2. Research the following types of commonly used cells and batteries, and identify concerns about their disposal:

- alkaline dry cell (zinc-manganese oxide)
- nickel-cadmium
- nickel-metal hydride
- lithium ion
- lead-acid

Prepare a poster, information pamphlet, podcast, or other form of communication to promote safer practices for the disposal of used cells and batteries in your local area.

Discuss

Place a copy of your poster, pamphlet, or presentation in the discussion area for your class. Review the presentations of other students in your class, and provide constructive feedback to your classmates in the following areas:

- Is the scientific information correct?
- Does the presentation address concerns about the disposal of different types of spent cells?
- Are effective communication strategies used? How might the presentation of information be improved?

Review the feedback you receive on your own presentation, and revise your presentation. Submit your final presentation to your teacher.

Module 4: Lesson 2 Assignment

Open the Assignment Booklet, and complete "Part 3" of "Module 4: Lesson 2 Assignment."

Lesson Summary

In Lesson 2 you considered the following question:

- How do you determine the cell potential for a voltaic cell?

You learned how to determine cell potential for a voltaic cell using the electrical potentials that correspond with the half-reactions listed in the "Table of Selected Standard Electrode Potentials." You completed an investigation to verify the predictions made for voltaic cells you constructed. You also investigated different types of consumer cells. You looked at the purpose of each type of cell and at concerns about their chemical components and proper disposal.

Module 4—Batteries and Balance

Lesson 3—Electrolytic Cells

Get Focused

Many of the devices you use daily rely on rechargeable electric cells. In this module you have learned about the ability of electric cells to convert chemical energy into electrical energy. What changes occur when the cell from your MP3 player or digital camera is "recharged"?

In this lesson you will learn about rechargeable cells. You will learn how a recharging cell is an example of a different type of electrochemical cell, and you will learn why some cells can be recharged while others cannot be recharged.

Consider the following questions as you complete Lesson 3:

- What is an electrolytic cell?
- What chemical and energy conversions occur within an electrolytic cell?

Module 4: Lesson 3 Assignment

In the Lesson 3 Assignment you will design and test an electrolytic cell. You will be directed to go to the Assignment Booklet to complete the Module 4: Lesson 3 Assignment later in this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

In previous science courses you may have seen how applying electrical energy to a solution caused chemical change. Electrolysis, the process of applying an electrical current to a system,

is used to cause chemical reactions. In Lesson 2 you learned that electric cells, voltaic cells for example, involve a spontaneous electrochemical change. What kind of change occurs in an electrolytic cell? Read pages 639–640 in the textbook to learn about the scientific principles behind the operation of an electrolytic cell.

You may wish to save a copy of "Table 1" from page 640 in your course folder as a reminder of the similarities and differences between the two types of electrochemical cells: electric and electrolytic.

To get a clear understanding of the differences between the behaviours of electric and electrolytic cells, read "Secondary Cells: Electric and Electrolytic" on page 640 of the textbook. Does it make sense that a rechargable cell can be described as both an electric and an electrolytic cell, but that the cell can't be both at the same time?

Self-Check

SC 1. Complete "Practice" questions 1–4 on page 640 of the textbook.

Check your work with the answer in the appendix.

Try This

Electrolysis of Potassium Iodide

In the next activity you will watch a video showing the electrolysis of a solution. The system you will observe is an aqueous potassium iodide solution that will be placed in a U-tube. A carbon electrode will be placed in each end of the tube. The carbon electrodes will be connected to a source of electricity.

Complete the following questions about the system before you view the video. The questions will prepare you to anticipate the events and changes you will observe in the video. As you complete the questions, recall the techniques that you used to predict redox reactions and that you applied in your study of voltaic cells. Remember to carefully scan the states of substances in half-reactions, as they may help you identify observable changes that will occur as the cell operates.

TR 1. What are the strongest oxidizing and reducing agents in this system?

TR 2. What are the half-reactions for the strongest oxidizing and reducing agents in this system?

TR 3. Describe one observable change you would expect to see at each of the electrodes, and describe a diagnostic test you could perform to confirm the result.

TR 4. Calculate the E°_{cell} for this system. What does this value represent in terms of operating the cell?

Watch and Listen

Go to the course multimedia DVD, and view the video "Electrolysis of Potassium Iodide." Record your observations and replay sections of the video to confirm your answers to questions TR 1–TR 3.

You may wish to read pages 641–642 in the textbook. This reading provides an in-depth explanation of the events observed in the video and will also confirm your answers to questions TR 1–TR 4.

Read

Read "Summary" on page 642 of the textbook. Work through "Communication example 1" and "Communication example 2" on pages 643 and 644.

Self-Check

SC 2. Complete "Practice" questions 5 and 6 on page 644 of the textbook.

Check your work with the answer in the appendix.

Self-Check

In previous science courses you may have seen the electrolysis of water that makes hydrogen and oxygen gases. In Module 3 you learned about disproportionation reactions, a type of electrochemical reaction in which the same species acts as the oxidizing and the reducing agent. The electrolysis of water is another example of a disproportionation reaction.

SC 3. What are the oxidation and reduction half-reactions involved in the electrolysis of water?

SC 4. What technical consideration is necessary to make the electrolysis of water possible?

Check your work with the answer in the appendix.

Self-Check

In a laboratory, the device used to collect the gases produced by an electrolysis is called a Hoffman's apparatus. The solution within the apparatus is electrolyzed, and the gases are collected in a chamber above each of the electrodes.

SC 5. Predict the products of the electrolysis of an aqueous sodium chloride solution using a Hoffman's apparatus.

Check your work with the answer in the appendix.

Watch and Listen

Go to the course multimedia DVD, and view the video "Electrolysis of Aqueous Sodium Chloride." Observe the Hoffman's apparatus for examples of change, especially in the parts of the cell above each of the electrodes. Do your observations confirm the predictions you made in your answer to SC 5?

Read

Read "Evaluation of Predicted Reactions—The Chloride Anomaly" on page 645 of the textbook to learn about the unique result obtained from the electrolysis of aqueous chloride solutions.

Read

What is the connection between electrolysis and the "recharging" that occurs within a rechargeable commercial cell? A rechargeable commercial cell does not allow new materials to enter or leave the cell. Therefore, the application of electrical energy used in recharging the cell must create a chemical change. The chemical change recreates the reactants of the spontaneous reaction that produces the flow of electrons. The spontaneous reaction that produces the flow of electrons is referred to as the discharge of the cell.

The reactions for the discharge of the cell and its recharging are reversible. The products of one process become the reactants for the next process. The process of discharging and recharging can be repeated many times, provided an energy source is available for the non-spontaneous recharging process, and that the half-reactions in the discharging process are reversible and recreate the reactants.

Reflect on the Big Picture

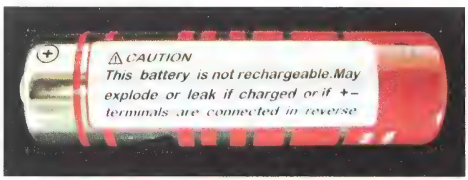
Use the information in the Read section above to complete the following questions. Use complete descriptions. Where possible, make reference to the information provided to support your answer.

RBP 1. The discharge of a Ni-Cd rechargeable cell is described by the following half-reactions:

$$2\ NiO(OH)(s) + 2\ H_2O(l) + 2\ e^- \rightarrow 2\ Ni(OH)_2(s) + 2\ OH^-(aq) \qquad E^\circ_{net} = +0.49\ V$$
$$Cd(s) + 2\ OH^-(aq) \rightarrow Cd(OH)_2(s) + 2\ e^- \qquad E^\circ_{net} = -0.81\ V$$

Arrange the half-reactions to determine the net reaction and the net cell potential for the discharge of a Ni-Cd cell. What are the products of the recharging process, and what is the minimum voltage that must be applied to a recharging Ni-Cd cell?

RBP 2. The half-reactions for a common type of dry cell are shown below:



The warning label reads "This battery is not rechargeable. May explode or leak if charged or it +– terminals are connected in reverse."

cathode: $MnO_2(s) + H_2O(l) + 2\ e^- \rightarrow Mn_2O_3(s) + 2\ OH^-(aq)$
anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-$

The cell shown in the photograph has a prominent warning label. Provide reasons for the warning shown in the picture, and explain why it is not physically possible to recharge this type of cell.

Save your answers in your course folder. Submit a copy of your answers to you teacher for feedback.

Module 4: Lesson 3 Assignment

Open the Assignment Booklet. Complete all questions in the "Module 4: Lesson 3 Assignment."

Lesson Summary

In Lesson 3 you considered the following questions:

- What is an electrolytic cell?
- What chemical and energy conversions occur within an electrolytic cell?

You found that although electrolytic and electric cells share many of the same characteristics, they have important differences. The table shown, from page 640 of the textbook, summarizes the similarities and differences between the voltaic cell (a type of electric cell) and the electrolytic cell.

Table 1—Comparing Electrochemical Cells: Voltaic [(a type of electric cell)] and Electrolytic

	Voltaic [(a type of electric)] cell	Electrolytic cell
spontaneity	spontaneous reaction	nonspontaneous reaction
standard cell potential E^{o}_{cell}	positive	negative
cathode	• strongest oxidizing agent present undergoes a *reduction* • positive electrode	• strongest oxidizing agent present undergoes a *reduction* • positive electrode
anode	• strongest reducing agent present undergoes an *oxidation* • negative electrode	• strongest reducing agent present undergoes an *oxidation* • positive electrode
direction of electron movement	anode → cathode	anode → cathode
direction of ion movement	• anions → anode • cations → cathode	• anions → anode • cations → cathode

In an electrolytic cell, electrical energy is applied to create chemical change. You learned that rechargeable cells, sometimes called secondary cells, are one type of electrolytic cell. Other electrolytic cells are used to produce gases like oxygen, hydrogen, and chlorine. In the next lesson you will learn more about how electrolytic cells are used to produce and refine metals.

Module 4—Batteries and Balance

Lesson 4—Commercial Applications of Electrolytic Cells

Get Focused

In Lesson 3 you saw how an electrolytic cell can be used to produce pure elements such as the gases hydrogen, oxygen, and chlorine, and pure metals like copper. The electrolytic cell causes chemical change by forcing a redox reaction to occur.

In previous science courses you may have seen a demonstration of the reactivity of sodium, potassium, and lithium (group 1 metals) with water. Do you recall which metals reacted most vigorously—so much so that you could see a flame? How could such reactive metals ever be converted into their elemental forms? Is it possible that an electrolytic cell could be used for the production of all metals?

In Lesson 3 you observed the operation of laboratory-scale electrolytic cells. How is the electrolytic cell used to plate chromium onto automobile bumpers, as shown in the photograph, similar and different to the cells you build in a laboratory? In this lesson you will investigate society's use of electrolytic cells as a technology.

Consider the following question as you complete Lesson 4:

- What are some of the practical applications of electrolytic cells?

Module 4: Lesson 4 Assignment

There is no assignment for this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

The development of electric cells was essential to experimentation using electrolytic cells. To learn more about this relationship and how electrolysis can be used to produce elements, read pages 646–647 in the textbook. Stop reading when you get to the end of "Communication example 3" on page 647.

The reading introduces an important aspect of the electrolysis of weak oxidizing agents like the ions of group I metals. Because water is a stronger oxidizing agent than $Na^+(aq)$, $K^+(aq)$, and $Li^+(aq)$ the production of metals in group I can only be achieved in molten conditions. You will recall that ionic compounds also tend to have high melting temperatures, normally well above the boiling point of water. When molten, ionic compounds such as NaCl(l) act as an electrolyte because ions are free to move. As you have already learned in this module, the movement of ions is essential for the conservation of charge in both electric and electrolytic cells.

Read "Production of Aluminium" on pages 647–648 in the textbook to see how an electrolytic cell containing molten contents is designed and operated. Then read "The Chlor-Alkali Process" on page 648 to learn how chlorine is produced at an industrial scale in a facility such as Dow Chemical in Fort Saskatchewan, Alberta.

In Unit C of this course you will learn how chlorine produced by the chlor-alkalai process can be used in chemical reactions with hydrocarbons to make plastics. The partnering of industrial processes like the chlor-alkali process with the refining of petroleum is an important part of Alberta's petrochemical industry.

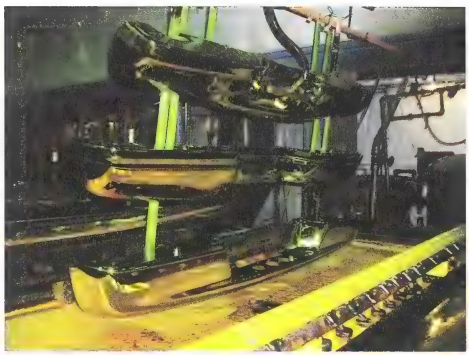
Self-Check

SC 1. Complete "Practice" questions 16–18 and 21 on page 649 of the textbook.

Check your work with the answer in the appendix.

Read

Electrolytic cells can also be used to refine metals. Read "Refining of Metals" on pages 649–650 of the textbook.



Many people have been attracted to the glowing chrome on classic and customized automobiles. The process of placing chromium over the steel used to construct bumpers and other parts of a car's body is called electroplating.

Watch and Listen

Go to the course multimedia DVD, and view the virtual investigation "Electroplating Copper." This investigation is a laboratory-scale example of electroplating.

TR 1. As you view the investigation, record qualitative and quantitative data observed.

Save a copy of the data in your course folder. You will use the data in Lesson 5 to complete an analysis of this investigation.

Try This

The following questions are based on the virtual investigation you just completed, "Electroplating Copper."

TR 2. Describe any changes that occurred at the carbon electrode.

TR 3. Describe any changes that occurred at the copper electrode.

TR 4. Identify the evidence that indicates a reduction half-reaction occurred during the operation of the electrolytic cell. Write the half-reaction that describes this reaction.

TR 5. Identify the evidence that indicates an oxidation half-reaction occurred during the operation of the electrolytic cell. Write the half-reaction that describes this reaction.

TR 6. Draw a diagram of the apparatus. Label the anode, cathode, electrolyte, and direction of electron flow provided by the power source.

Read

Read "Electroplating" on page 650 of the textbook.

Self-Check

SC 2. Complete "Section 14.3" question 15 on page 651 of the textbook.

Check your work with the answer in the appendix.

Reflect on the Big Picture

Society uses a lot of metal. That makes the industries that use electrolytic cell technology to produce elements and refine and plate metals extremely important. There are toxicity issues associated with some metals and substances used in electroplating. Careful treatment of chemical waste is essential if these industries are to demonstrate sustainability.

You may wish to investigate whether any businesses in your local area electroplate or recycle metals (recycling also often uses electroplating). How do these businesses deal with chemical waste? What provincial and federal laws apply to these businesses? How do these businesses address concerns about protecting the environment?

Lesson Summary

In Lesson 4 you considered the following question:

- What are some of the practical applications of electrolytic cells?

You learned that electrolytic cells are used to produce elements, refine metals, and electroplate. In the next lesson you will learn about quantitative relationships that need to be considered when operating an electrolytic cell. Can you think of what chemical entity in an electrolytic cell determines the quantity of product that can be produced? Go to Lesson 5 to learn the answer.

Module 4—Batteries and Balance

Lesson 5—Quantitative Relationships in Cells

Get Focused

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In Lesson 4 you learned that electrolytic cells can be used to produce measurable quantities of elements, including gases and electroplated metals. As you completed the virtual investigation "Electroplating Copper," did you find yourself wondering about the changes to the mass of each electrode in the cell? In this lesson you will revisit the data you collected from the Lesson 4 virtual investigation, and you will research other electrolytic cells. You will investigate the stoichiometric relationships involved.

From your work in previous chemistry courses and in earlier modules of this course, you will recall that stoichiometry is a process that can be used to predict and explain the chemical and measurable quantities of matter involved in a process. Can you think of how stoichiometric principles could be applied to an electrochemical reaction occurring in an electric or electrolytic cell? Is it possible to have a quantitative relationship between electrons and the atoms involved?

Consider the following question as you complete Lesson 5:

- What are the important quantitative relationships within operating electric and electrolytic cells?

Module 4: Lesson 5 Assignment

You will be directed to go to the Assignment Booklet to complete the Module 4: Lesson 4 Assignment later in this lesson.

You must decide what to do with the questions that are not marked by the teacher.

Remember that these questions provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

Earlier in this module you were asked to list similarities and differences between electric and electrolytic cells. One important similarity is that both cell types need a transfer of electrons in order for any chemical change to occur. You tested this when you constructed a voltaic cell and observed the change in the voltmeter reading when the cell's circuit was broken by briefly removing various components. In an electrolytic cell, you observed that a chemical change only occurs when the power source is turned on and a voltage greater than the minimum cell potential is applied. Is it possible to identify what chemical entity controls the change observed in other species? What do these observations tell you about electrochemical cells in terms of quantitative relationships?

You probably haven't considered electrons as quantifiable. However, when you wrote half-reactions, you used coefficients to indicate the proportions of electrons involved in an oxidation or reduction. Read the text and work through the "Sample problems" and "Communication examples" on pages 652–654 in the textbook to learn more about the methods used to quantify electrons.

Self-Check

SC 1. Complete "Practice" questions 1–4 on page 653 of the textbook and "Practice" questions 5–7 on page 654 of the textbook.

Check your work with the answer in the appendix.

Read

In the previous section you learned how to use measurements of amperage and time to calculate moles of electrons. As you will recall, a balanced chemical reaction, including a half-reaction, provides information about the proportions of substances in a chemical process. With knowledge of the quantity of electrons, you will be able to complete the same kinds of stoichiometric calculations performed elsewhere in this course to determine the quantities of other species in the half-reaction.

Read pages 655–656 in the textbook. Work through "Sample problem 14.4" and "Communication example 3" to learn more about performing these calculations.

Self-Check

SC 2. Complete "Section 14.4" questions 1–3, 9, 11, and 12 on page 657 of the textbook.

Check your work with the answer in the appendix.

Module 4: Lesson 5 Assignment

Open the Assignment Booklet. Complete "Module 4: Lesson 5 Assignment."

Try This

Earlier in this module you investigated voltaic cells. Could the calculations involving amperage, time, and moles of electrons be applied to predicting the changes involved in the operation of a voltaic cell?

TR 1. Demonstrate the application of stoichiometry to predicting a quantitative relationship in a voltaic cell of your choice. The cell can be a cell described in the textbook or in other resources, or a cell with reactants of your choice. Your demonstration can be presented in a format similar to the "Sample problems" or "Communication examples" in the textbook.

Save a copy of your answer in your course folder. Submit a copy of your answer to your teacher for feedback.

Reflect on the Big Picture

In this lesson you learned how to use quantitative information related to the operation of an electrochemical cell to predict the amount of species participating in the reaction. Throughout this module you have considered the environmental impact of electrochemical processes. Like other chemical processes, the properties and quantity of chemical waste are important in determining appropriate treatment methods.

Future research in the area of environmental chemistry or in other scientific areas may change opinions regarding the harm that some forms of chemical waste may cause. For example, alkaline dry cells collected at Eco Station facilities in Edmonton are currently placed in a special landfill; depending on future developments in technology for recycling, these alkaline dry cells may be unearthed and processed. It is very possible that the processing of waste from alkaline dry cells will involve further electrochemical reactions, possibly even electrolysis. Therefore, a knowledge of quantitative relationships in electrochemical processes is certain to be an important part of any efforts to recycle this type of waste.

Lesson Summary

In Lesson 5 you considered the following question:

- What are the important quantitative relationships within operating electric and electrolytic cells?

This lesson extended your application of the stoichiometric method to include amperage and time. You used the stoichiometric method to determine the chemical quantity of electrons participating in a reaction. Since electrons determine the chemical changes in these cells, they can be treated as the limiting reactant in many electrochemical systems. As such, electrons can be used to calculate the quantities of products or reactants in the system that participates in the process.

Module 4—Batteries and Balance

Module Summary

In this module you investigated technological applications of the scientific principles associated with reduction–oxidation reactions.

You considered the following module question:

- How are principles of oxidation and reduction applied to electrochemical cells?

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You used your understanding of electron transfer reactions and skills and techniques to predict, explain, and interpret chemical changes in order to investigate electric (e.g., voltaic) and electrolytic cells.

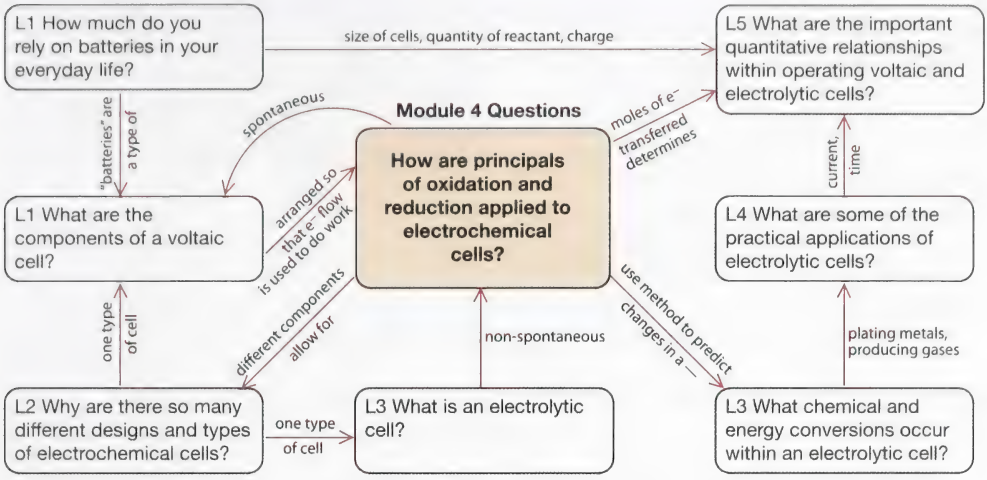
As you found with the chemical systems you investigated in Module 3, the chemical behaviour of both electric and electrolytic cells can be explained using half-reactions, net ionic equations, and cell potentials. You paid careful attention to the products and reactants in the equations you wrote. This allowed you to make observations that confirmed your predictions for both types of electrochemical cells.

You also learned to apply your knowledge of stoichiometry to reduction-oxidation reactions occurring in these cells. You extended your understanding of reaction chemical entities by using information about amperage and time to calculate moles of electrons participating in a chemical process. In the case of an electrolytic cell, you regarded electrons as the limiting reagent in the cell's operation.

Concept Map or Graphic Organizer

As you worked through Module 4, you may have added information to a concept map or graphic organizer based on the module and lesson questions listed in the Module 4 Concept Organizer. Now is a good time to review the relationships in your concept map or graphic organizer and to try to answer the module and lesson questions.

The sample Module 4 concept map shown below provides one set of possible links between the questions. Remember that this is one possible description only—there are many other correct possibilities. However, if your completed concept map or graphic organizer differs significantly from the sample, you may wish to contact your teacher or to compare your map or organizer with those of other students in your class. This will ensure that your interpretations of lesson materials and your descriptions are accurate.


L1 How much do you rely on batteries in your everyday life?
size of cells, quantity of reactant, charge
L5 What are the important quantitative relationships within operating voltaic and electrolytic cells?
"batteries" are a type of
spontaneous
Module 4 Questions
moles of e⁻ transferred determines
current, time
L1 What are the components of a voltaic cell?
arranged so that e⁻ flow is used to do work
How are principals of oxidation and reduction applied to electrochemical cells?
L4 What are some of the practical applications of electrolytic cells?
one type of cell
different components allow for
non-spontaneous
use method to predict changes in a
plating metals, producing gases
L2 Why are there so many different designs and types of electrochemical cells?
one type of cell
L3 What is an electrolytic cell?
L3 What chemical and energy conversions occur within an electrolytic cell?

Module Assessment

When you were in elementary school, you might have constructed an electrochemical cell using a potato or a lemon. You have come a long way in understanding electrochemical cells. In this Assessment you will demonstrate your understanding of the scientific principles and technical knowledge involved in the design and operation of electrochemical cells.

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With an electrochemical cell made from a potato or a lemon in mind, investigate ONE of the following research questions:

- Do all cells of this kind produce the same voltage?
- Do all cells of this kind operate for the same length of time?
- Can different voltages be recorded from cells constructed using the same fruit or vegetable?
- Is it possible to make these kinds of cells using a wide range of fruits and vegetables?

Read through the parts of conducting an experiment described in the table provided. Formally document your work in two of these steps, and submit your documentation to your teacher for assessment. Document your work in the format that best allows you to provide accurate, detailed, and complete descriptions of your work; e.g., video, audio, text, or a combination of formats.

Part	Criteria
Planning an Experiment	• State the problem to be investigated. • List manipulated, responding, and controlled variables for the proposed experiment. • Describe an experiment to test the problem developed. • Describe how variables will be manipulated or controlled. • Provide instructions about what observations and data must be collected, including experimental controls. • List conclusions that cannot be made from the proposed experiment (identify design limitations and possible conclusions).
Performing an Experiment	• Provide evidence of the student's own experiment. • Write the procedure describing the experiment that was carried out, and describe how the data were collected. • Include raw data and other observations for all trials and controls. • Include organized (processed) data from the experiment.
Analysis of Experimental Work/Design	• Thoroughly describe the processes used to analyze the data. • State a conclusion that can be supported by the data analyzed. • List conclusions that cannot be made from the proposed experiment (identify design limitations). • Identify a new question or problem that arises from this experiment.

Unit B Conclusion

In Unit B you investigated reduction-oxidation reactions, and you learned about technologies that apply these kinds of reactions.

In Module 3 you learned about the principles involved in electrochemical change. You investigated the following module questions:

- What properties of metals make them popular choices in the construction and production of materials?

- How can an understanding of corrosion allow for better selection of materials and development of methods that reduce material damage?

You interpreted empirical evidence from a variety of chemical systems. As you completed investigations, you saw evidence of chemical reactions including colour change, formation of gases, and formation of precipitates. You learned to use observations to discriminate between spontaneous and non-spontaneous processes and to predict their occurrence using the "Table or Selected Standard Electrode Potentials." You described the chemical changes you observed using half-reactions and net ionic chemical equations.

These skills enabled you to identify instances in which corrosion of metals occurs, and to predict situations that would encourage the chemical process of corrosion. Finally, you learned how redox reactions can be used to perform quantitative analysis.

In Module 4 you investigated the following module question:

- How are principles of oxidation and reduction applied to electrochemical cells?

You applied the skills you learned in Module 3 to learn about the electrochemical cell. You used your ability to predict and describe electrochemical changes in the design of electric (voltaic) and electrolytic cells, and you used you observation skills and basic chemical tests to confirm the chemical changes observed at the electrodes of these cells.

You continued to advance you understanding of quantitative relationships in electrochemical systems by learning to consider the quantity of electrons as a chemical entity and to perform stoichiometric calculations involving current, time, and moles of chemical substances. Finally, you learned about other types of electrochemical cells (including fuel cells and electrolytic cells) and their application as technologies used by society.

You should now be able to

- explain the nature of oxidation-reduction reactions
- apply the principles of oxidation-reduction to electrochemical cells
- analyze scientific and technological systems that apply electrochemical principles

Your study of electrochemical changes is an important part of understanding chemical change, a central theme of Chemistry 30. In unit D you will be introduced to the concept of proton-transfer reactions. In Unit D and as a result of your learning in Unit B, you will be able to apply many principles into another area of chemistry—chemical equilibrium.

Self-Check

Go to the course multimedia DVD, and complete the "Unit B Diagnostic Self-Check" to review the concepts, skills, and knowledge introduced and to assess your understanding of them.

If you have trouble answering any of the questions, try one or more of the following options:

- Go to the lesson(s) indicated for each question. Once there, review the information provided and the relevant sections of the textbook.
- View the hint associated with each question.
- Look at the answer and then work backward to understand the problem and its solution.
- Ask your teacher for help.

You understand how important it is to review material in preparation for any tests your teacher may give you and in solidifying your understanding of new concepts. Completing these questions is just one part of your review. Program of Studies information is associated with each question. You may wish to use this information to identify areas in which you need to pay particular attention. You may also answer the questions that appear at the end of each chapter and unit in the textbook.

You have already reviewed your Modules 3 and 4 concept maps or graphic organizers. Now is a good time to give them another look and add any new connections or ideas to them.

Unit Assessment

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In Unit B you learned about electrochemical change. The contents of an emergency kit are pictured. Choose one item from the kit. Using resources like the textbook, the Internet, and product information, determine how the operation of the chosen device involves an electrochemical change.

Use relevant chemical equations and terms to explain the electrochemical change that occurs and to explain how this electrochemical change contributes to the function of the device. State one positive or one negative aspect of the device's use that relate to the electrochemical change involved in its operation. You may wish to support your response with graphics or an audio recording.

Developing Your Response

The format for the Unit B Assessment and its response is intended to familiarize you with the expectations of other assessments you may encounter in this course, including unit tests administered by your teacher and the Chemistry 30 Diploma Examination. Pay careful attention to the question elements and to the descriptions of performance—these can improve your ability to address these kinds of questions and your test results.

When answering detailed questions like this one, it's important to review and understand the criteria that will be used to assess your response. Go to the course multimedia DVD, and open "Making Sense of Scoring Criteria." Use "Making Sense of Scoring Criteria" and the scoring guide provided below to plan and develop your response.

It is important to review and understand the criteria in a scoring guide before you start to develop your response, since the scoring guide gives you all of the points you need to address in your answer. Consider the descriptive terms used in the scoring guide to make sure you properly develop ideas and use examples and other information to support your ideas. You should also use the criteria to when you proofread and revise your response before submitting your response to your teacher.

Save your completed response in your course folder, and submit a copy to your teacher.

Score	Scoring Description
5 **Excellent**	The response demonstrates an **excellent** understanding of an electrochemical change occurring in one component of the emergency kit and uses a **correct chemical equation** and an **accurate and thorough explanation** of the electrochemical change involved. The explanation of the contribution of the electrochemical change to the overall function of the device is **explicit** and **well thought-out**. Statements of benefits and negative aspects are **accurate and insightful**.
4 **Good**	The response demonstrates a **good** understanding of an electrochemical change occurring in one component of the emergency kit and uses a **correct chemical equation** and an **accurate explanation** of the electrochemical change involved. The explanation of the contribution of the electrochemical change to the overall function of the device is **correct** and **well thought-out**. Statements of benefits and negative aspects are **accurate and insightful**.
3 **Satisfactory**	The response demonstrates a **satisfactory** understanding of an electrochemical change occurring in one component of the emergency kit and uses a **correct chemical equation** and a **plausible explanation** of the electrochemical change involved. The explanation of the contribution of the electrochemical change to the overall function of the device is **generally correct**. Statements of benefits and negative aspects are **correct**.
2 **Limited**	The response demonstrates a **poor** understanding of an electrochemical change occurring in one component of the emergency kit and uses a **plausible chemical equation.** Accompanying explanations are sketchy and lack development. The explanation of the contribution of the electrochemical change to the overall function of the device is **vague**. Benefits and negative aspects are **listed**.
1 **Poor**	The student provides a solution that contains a relevant statement that begins to answer the question.
0 **Insufficient**	The response is incorrect and/or totally off topic.

Appendix

Module 4 Self-Check Answers

Contact your teacher if your answers vary significantly from the answers provided here.

Lesson 1

SC 1. Copper and zinc are used in the cell. Zinc is the more reactive metal. The copper ion is the more reactive metal ion.

SC 2. The redox half-reactions will involve the reduction of the most reactive metal ion and the oxidation of the most reactive metal.

reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

SC 3. The reaction will be spontaneous because the reduction half-reaction appears above the oxidation half-reaction in the "Table of Selected Standard Electrode Potentials." From experiments performed in Module 3, it is known that reactants with similar positions on the table react spontaneously.

SC 4. As stated in the half-reactions, the zinc will lose electrons and the copper ions will gain electrons.

SC 5. The zinc electrode should reduce in mass and possibly in size as the zinc atoms lose electrons and are changed into zinc ions that will go into solution. The copper electrode should increase in mass as the copper ions come out of the solution to join with electrons to form copper metal.

SC 6. A voltaic cell is a closed chemical system, and the reactants will be depleted as the reactions proceed. Unless more copper(II) ions are placed into the cell and more zinc is added, no reactants will be present to change electrons.

SC 7. Since the zinc metal is the site of oxidation and, therefore, the source of electrons, it is the negative electrode. The copper electrode is the destination for the electrons in contact with the copper ions that will gain the electrons. This makes the copper metal the positive electrode.

SC 8. When the salt bridge was lifted out of the solution, the voltage dropped to zero.

SC 9. Hypothesis: Systems need to maintain a balance in their function. In this cell, electrons transfer between the parts of the cell. When electrons transfer, there is a flow of negative charge. To maintain a balance, there must also be a flow of positive charge to balance the flow of negative charge.

The salt bridge allows ions in the electrolytes to move, thus completing the flow of charge in the

circuit. If electrons are flowing along the wire from one electrode to the other, there must be a movement of ions within the cell to complete the circuit. This is completed by the movement of dissociated ions in the solutions through the salt bridge.

SC 10. When the leads from the voltmeter were connected to the opposite electrodes in the voltaic cell, the voltmeter switched from displaying a positive output to displaying a negative output. If the original output was negative, it would have switched to positive. In both cases, the output changes because the electrons are flowing through the voltmeter in the opposite direction.

SC 11. The copper electrode had a reddish-brown sediment on its surface. This is the copper metal being deposited from the copper ions that come out of the solution as described by the reduction half-reaction. Since the copper ions are being removed from the solution containing the copper electrode, this solution appeared to have a slightly less intense colour after the cell had been operating. Although no visible changes could be observed on the zinc electrode, this electrode would diminish in mass and size if the voltaic cell were able to run for a long time.

SC 12.

Practice 1.

voltaic cell: a device consisting of two half-cells connected by an external circuit and a porous boundary or salt bridge

half-cell: a portion of a voltaic cell consisting of an electrode and an electrolyte

porous boundary: a separation between the two half-cells in a voltaic cell; this boundary must allow for the exchange of ions dissolved in solution between the half-cells

salt bridge: one type of porous boundary

electrolyte: the solution within a half-cell or within the salt bridge

external circuit: the wire connecting the two electrodes

inert electrode: an electrical conductor such as carbon or platinum that is not a reactant in the electrochemical reaction occurring within the cell

Practice 2.

cathode: the electrode in which the reduction reaction involving the strongest oxidizing agent in the system occurs

anode: the electrode in which the oxidation reaction involving the strongest reducing agent in the system occurs

Practice 3.

a. cathode
b. anode
c. anode
d. cathode

Practice 4.

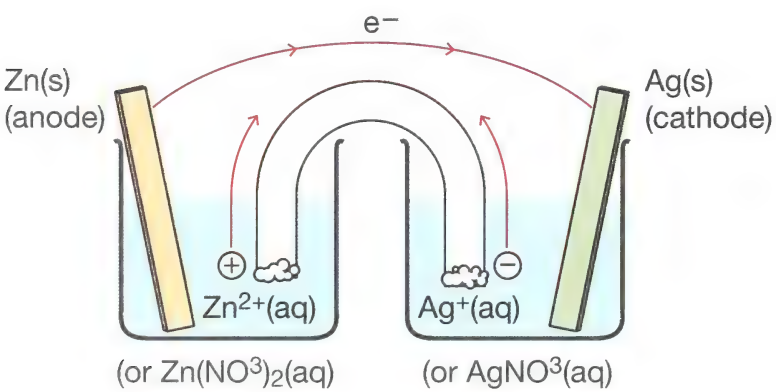
An inert electrode is used when the strongest oxidizing or reducing agent in a voltaic cell is in the aqueous or molten state. The inert electrode allows the transfer of electrons in a half-cell. Graphite (carbon) and platinum are two commonly used inert electrodes.

Practice 5.

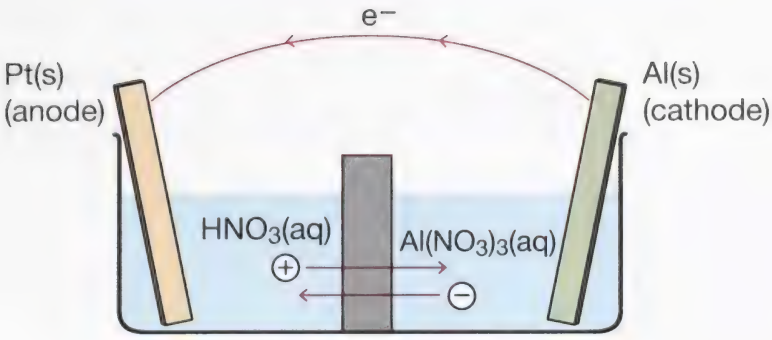
The solute in the salt bridge solution should be a very soluble ionic compound containing spectator ions (ions that will not participate in a reaction within the cell).

Practice 6.

a. strongest oxidizing agent: $Ag^+(aq)$
strongest reducing agent: $Zn(s)$
cathode half-reaction: $(Ag^+(aq) + 1e^- \rightarrow Ag(s)) \times 2$
anode half-reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
net cell reaction: $2\ Ag^+(aq) + Zn(s) \rightarrow 2\ Ag(s) + Zn^{2+}(aq)$



b. strongest oxidizing agent: $NO_3^-(aq)$, $H^+(aq)$
strongest reducing agent: $Al(s)$
cathode half-reaction: $(2\ NO_3^-(aq) + 4\ H^+(aq) + 2e^- \rightarrow N_2O_4(g) + 2\ H_2O(l)) \times 3$
anode half-reaction: $(Al(s) \rightarrow Al^{3+}(aq) + 3e^-) \times 2$
net cell reaction: $6\ NO_3^-(aq) + 12\ H^+(aq) + 2\ Al(s) \rightarrow 3\ N_2O_4(g) + 6\ H_2O(l) + 2\ Al^{3+}(aq)$



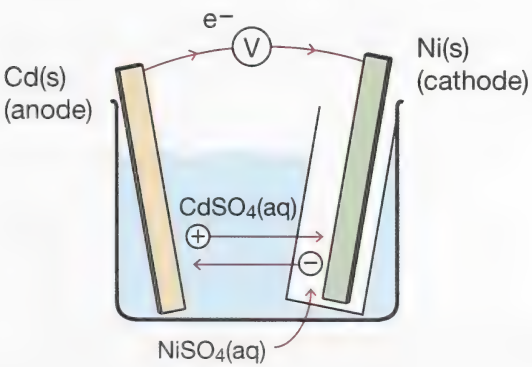
Practice 7.

a. Cations migrate toward the cathode; anions migrate toward the anode.

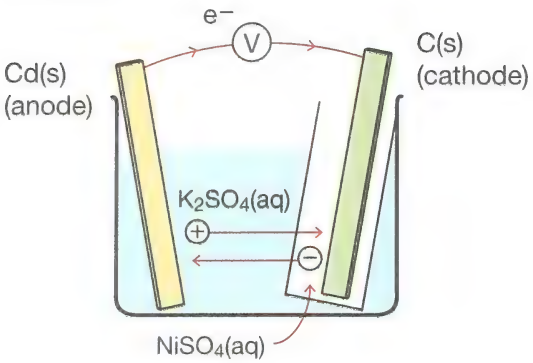
b. Because electrons are moving from the anode half-cell to the cathode half-cell, ions must move to keep all parts of the cell electrically neutral. This raises the question of why the electrons move in the first place. The answer is that the electrodes have different potentials. The anode has a greater tendency to lose electrons than does the cathode.

c. In a cell, all positive ions (cations) migrate toward the cathode. Since Cu^{2+} is a blue ion, its colour becomes visible in the salt bridge as the ions pass through from anode to cathode.

Practice 8.



Practice 9.



The net reaction for the cells in Practice 8 and Practice 9 is $Cd(s) + Ni^{2+}(aq) \rightarrow Cd^{2+}(aq) + Ni(s)$.

Lesson 2

SC 1.

	Half-Reaction
Reduction	$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$
Oxidation	$Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-$

$$E^\circ_{cell} = +0.34\ V - (-0.76\ V)$$
$$= +1.10\ V$$

SC 2.

	Half-Reaction
Reduction	$Ag^+(aq) + 1\ e^- \rightarrow Ag(s)$
Oxidation	$Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^-$

$$E^\circ_{cell} = +0.80\ V - (+0.34\ V)$$
$$= +0.46\ V$$

SC 3. The separation between the copper and zinc half-reactions is greater than the separation between the copper and silver half-reactions. This is consistent with the differences in cell potentials observed. Therefore, the observation is supported by the examples analyzed.

SC 4. The chemicals used in the three types of cells is the same, therefore the cell potentials should be identical.

SC 5.

Practice 12.

a. $Pb(s) \mid Pb^{2+}(aq) \mid\mid Cu^{2+}(aq) \mid Cu(s)$

The cathode is Cu(s); the anode is Pb(s).

$$E^0_{cell} = +0.34\ V - (-0.13\ V)$$
$$= +0.47\ V$$

b. $Zn(s) \mid Zn^{2+}(aq) \mid\mid Ni^{2+}(aq) \mid Ni(s)$

The cathode is Ni(s); the anode is Zn(s).

$$E^0_{cell} = -0.26\ V - (-0.76\ V)$$
$$= +0.50\ V$$

c. $Pt(s) \mid H_2(g), H^+(aq) \mid\mid Fe^{3+}(aq), Fe^{2+}(aq) \mid Pt(s)$

The cathode and anode are both Pt(s).

$$E^0_{cell} = +0.77\ V - 0.00\ V$$
$$= +0.77\ V$$

Note that H_2 can reduce Fe^{3+} to Fe^{2+} but can't reduce it all the way to Fe.

Practice 13.

$E^\circ_{cell} = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode})$

$E^\circ_r(\text{anode}) = E^\circ_r(\text{cathode}) - E^\circ_{cell}$

$$E^0_r(\text{anode}) = +1.50\ V - (+1.84\ V)$$
$$= -0.34\ V$$

The reduction potential of indium(III) ion is –0.34 V.

Practice 14.

For the copper half-cell, E°_{r} would become +3.38 V.

For the zinc half-cell, E°_{r} would become +2.28 V.

In the shift to this new reference point, all the reduction potentials must increase by the same amount (3.04 V) so that the differences between individual values remain constant.

Practice 15.

Differences could be caused by non-standard conditions of temperature, pressure, or concentration. Other possible factors include purity of substances, oxide coating on electrodes, and variations in the salt bridge or porous boundary.

Practice 16.

The cell "goes dead" because the reaction $Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$ has proceeded to the point of equilibrium. At this point there is a high concentration of $Zn^{2+}(aq)$ and a low concentration of $Fe^{2+}(aq)$.

Because you have not yet studied Module 7 on equilibrium, your answers may not be as complete as those provided here.

Lesson 3

SC 1.

Practice 1. At the cathode of an electrolytic cell, the oxidizing agent gains electrons and is reduced. At the anode of the cell, the reducing agent loses electrons and is oxidized.

Practice 2. In a voltaic cell, the cathode and anode are designated positive and negative respectively. In an electrolytic cell, the designations are the opposite.

Practice 3. In an electrolytic cell, electrons move through the external circuit from anode to cathode. Cations migrate toward the cathode, and anions migrate toward the anode.

Practice 4. The pumping of water into the tower is similar to the recharging of a secondary cell because energy is applied to drive a change (by applying a current to force a chemical reaction to occur in the cell). When water flows from the tower, it does so due to gravity. This spontaneous process is similar to the discharging of the secondary cell in which the chemical energy is converted into electrical energy.

SC 2.

Practice 5.

a. cathode: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$

anode: $2\ I^-(aq) \rightarrow I_2(aq) + 2e$

net cell: $Ni^{2+}(aq) + 2\ I^-(aq) \rightarrow Ni(s) + I_2(aq)$

$$E^0_{cell} = -0.26\ V - (+0.54\ V)$$
$$= -0.80\ V$$

A minimum voltage of +0.80 V is required.

b. cathode: $(2\ H_2O(l) + 2e^- \rightarrow H_2(g) + 2\ OH^-(aq)) \times 2$

anode: $4\ OH^-(aq) \rightarrow O_2(g) + 2\ H_2O(l) + 4e^-$

net cell: $2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$

$$E^0_{cell} = -0.83\ V - (+0.40\ V)$$
$$= -1.23\ V$$

A minimum voltage of +1.23 V is required.

Practice 6.

a. $$E^0_{cell} = -0.41\ V - (+1.07\ V)$$
$$= -1.48\ V$$

The minimum voltage required is +1.48 V. Note that the oxidizing agent is Cr^{3+} and the reducing agent is Br^-.

b. $$E^0_{cell} = +0.34\ V - (+1.23\ V)$$
$$= -0.89\ V$$

The minimum voltage required is +0.89 V. Note that the oxidizing agent is Cu^{2+} and the reducing agent is H_2O.

SC 3.

reduction (cathode): $2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$

oxidation (anode): $2\ H_2O(l) \rightarrow O_2(g) + 4\ H^+(aq) + 4\ e^-$

SC 4. Pure water does not conduct an electric current; therefore, a flow of charge would not be possible within an electrolytic cell containing pure water. Often the water has a small quantity of sodium sulfate added to it prior to electrolysis. The sodium ions and the sulfate ions in the system are spectator ions, but they are essential to allowing the cell to complete the flow of charge necessary for the transfer of electrons to occur.

SC 5. The strongest oxidizing and reducing agent in this system is water; therefore, the expected products are as follows:

- hydrogen gas and hydroxide ions (cathode; basic pH)
- oxygen gas and hydrogen ions (anode; acidic pH)

Lesson 4

SC 1.

Practice 16.

From an aqueous solution, trying to reduce a metal ion (which is a weaker oxidizing agent than water) is futile. When water is the strongest oxidizing agent present, it will be reduced and the metal ion will remain unchanged in the solution. Also, many ionic compounds have low solubility in water, which makes it impractical to electrolyze an aqueous solution to obtain a metal. The technological solution is to conduct the electrolysis in the absence of water; which is possible by melting the ionic compound.

Practice 17.

cathode: $2[Sc^{3+}(l) + 3\ e^- \rightarrow Sc(s)]$
anode: $3[2\ Cl^-(l) \rightarrow Cl_2(g) + 2\ e^-]$
net cell: $2\ Sc^{3+}(l) + 3\ Cl^-(l) \rightarrow 2\ Sc(s) + 3\ Cl_2(g)$

Practice 18.

cathode: $2\ H_2O(l) + 2e^- \rightarrow H_2(g) + 2\ OH^-(aq)$
anode: $2\ Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
net cell: $2\ H_2O(l) + 2\ Cl^-(aq) \rightarrow H_2(g) + 2\ OH^-(aq) + Cl_2(g)$

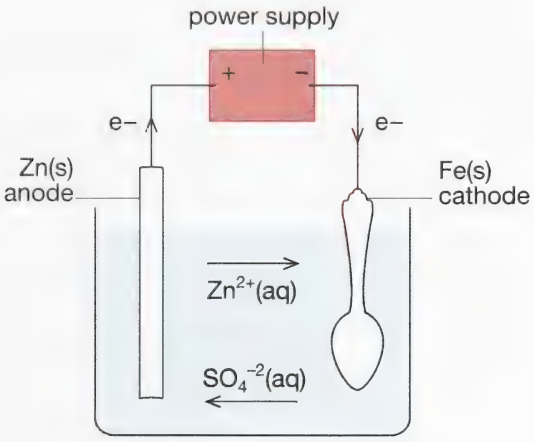
Practice 21.

Recycling aluminium has many benefits. To recycle an aluminium can, it takes only 5% of the energy required to produce a new one. Because every method for producing electrical energy has negative consequences, this energy saving reduces the aluminium industry's impact on the environment. Recycling aluminium also saves bauxite, the mineral from which aluminium is obtained. This preserves bauxite for future generations. Recycling also reduces the amount of waste going into landfills.

SC 2.

Section 14.3 15.

Electroplating Zinc onto an Iron Spoon



Lesson 5

SC 1.

Practice 1.

$$\begin{aligned} Q &= It \\ &= (1.5 \text{ C/s})(30 \text{ s}) \\ &= 45 \text{ C} \end{aligned}$$

The charge transferred is 45 C.

Practice 2.

$$I = \frac{Q}{t}$$
$$= \frac{(87.6\ \text{C})}{(22.5\ \text{s})}$$
$$= 3.89\ \text{C/s}$$

The electric current is 3.89 C/s.

Practice 3.

$$Q = It$$
$$= \left(250 \times 10^{-3}\ \text{C/s}\right)(28.5\ \text{s})$$
$$= 7.13\ \text{C}$$

The charge transferred is 7.13 C.

Practice 4.

$$t = \frac{Q}{I}$$
$$= \frac{(375\ \text{C})}{(1.60\ \text{C/s})}$$
$$= 234\ \text{s}$$

$$t = (234\ \text{s}) \times \left(\frac{1\ \text{min}}{60\ \text{s}}\right)$$
$$= 3.91\ \text{min}$$

The time taken is 3.91 min.

Practice 5.

$$t = 35\ \text{min} \times \left(\frac{60\ \text{s}}{1\ \text{min}}\right)$$
$$= 2.1 \times 10^{3}\ \text{s}$$

$$n_{\text{e}} = \frac{It}{f}$$
$$= \frac{(1.9\ \text{C/s})\left(2.1 \times 10^{3}\ \text{s}\right)}{\left(9.65 \times 10^{4}\ \text{C/mol}\right)}$$
$$= 0.041\ \text{mol}$$

The amount of electrons transferred is 0.041 mol.

Practice 6.

$$t = \frac{n_e F}{I}$$

$$= \frac{(0.146 \text{ mol})\left(9.65 \times 10^4 \text{ C/mol}\right)}{(1.24 \text{ C/s})}$$

$$= 1.14 \times 10^4 \text{ s}$$

$$t = \left(1.14 \times 10^4 \text{ s}\right) \times \left(\frac{1 \text{ h}}{3600 \text{ s}}\right)$$

$$= 3.16 \text{ h}$$

The transfer time was 3.16 h.

Practice 7.

$$t = (20 \text{ min}) \times \left(\frac{60 \text{ s}}{1 \text{ min}}\right)$$

$$= 1.2 \times 10^3 \text{ s}$$

$$I = \frac{n_e F}{t}$$

$$= \frac{(0.015 \text{ mol})\left(9.65 \times 10^4 \text{ C/mol}\right)}{\left(1.2 \times 10^3 \text{ s}\right)}$$

$$= 1.2 \text{ C/s}$$

The current required is 1.2 A.

SC 2.

Section 14.4 1.

$$t = (15.0 \text{ min}) \times \left(\frac{60 \text{ s}}{1 \text{ min}}\right)$$
$$= 900 \text{ s}$$

$$n_e = \frac{It}{F}$$
$$= \frac{(0.300 \text{ C/s})(900 \text{ s})}{(9.65 \times 10^4 \text{ C/mol})}$$
$$= 0.002\,80 \text{ mol}$$

The amount of electrons transferred is 0.002 80 mol.

Section 14.4 2.

$2\ Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

$$t = (8.0 \text{ h})\left(\frac{3600 \text{ s}}{1 \text{ h}}\right)$$
$$= 2.9 \times 10^4 \text{ s}$$

$$n_e = \frac{It}{F}$$
$$= \frac{(55 \times 10^3 \text{ C/s})(2.9 \times 10^4 \text{ s})}{(9.65 \times 10^4 \text{ C/mol})}$$
$$= 1.6 \times 10^4 \text{ mol}$$

$$n_{Cl_2} = 1.6 \times 10^4 \text{ mol} \times \frac{1}{2}$$
$$= 8.2 \times 10^3 \text{ mol}$$

$$m_{Cl_2} = (8.2 \times 10^3 \text{ mol})(70.90 \text{ g/mol})$$
$$= 5.8 \times 10^5 \text{ g}$$

The mass of chlorine formed is 5.8 x 10^5 g, or 5.8 × 10^2 kg.

Section 14.4 3.

$$Ag^{+}(aq) + 1e^{-} \rightarrow Ag(s)$$

$$m_{Ag} = \left(\frac{10.00\ \text{g}}{107.87\ \text{g/mol}}\right)$$
$$= 0.092\ 70\ \text{mol}$$

$$n_e = 0.092\ 70\ \text{mol} \times \frac{1}{1}$$
$$= 0.092\ 70\ \text{mol}$$

$$t = \frac{n_e F}{I}$$
$$= \frac{(0.092\ 70\ \text{mol})\left(9.65 \times 10^4\ \text{C/mol}\right)}{(1.80\ \text{C/s})}$$
$$= 4.97 \times 10^3\ \text{s}$$

$$t = \left(4.97 \times 10^3\ \text{s}\right) \times \left(\frac{1\ \text{min}}{60\ \text{s}}\right)$$
$$= 82.8\ \text{min}$$

The time required is 82.8 min.

Section 14.4 9.

a. $Cr_3^+(aq) + 3e^- \rightarrow Cr(s)$

$$t = (45.5 \text{ min}) \times \left(\frac{60 \text{ s}}{1 \text{ min}}\right)$$
$$= 2.7 \times 10^3 \text{ s}$$

$$n_e = \frac{It}{F}$$
$$= \frac{(54 \text{ C/s})\left(2.7 \times 10^3 \text{ s}\right)}{\left(9.65 \times 10^4 \text{ C/mol}\right)}$$
$$= 1.5 \text{ mol}$$

$$n_{Cr} = 1.5 \text{ mol} \times \frac{1}{3}$$
$$= 0.51 \text{ mol}$$

$$m_{Cr} = (0.51 \text{ mol})(52.00 \text{ g/mol})$$
$$= 26 \text{ g}$$

The mass of chromium deposited is 26 g.

b. $Ni_2^+(aq) + 2e^- \rightarrow Ni(s)$

$$n_{Ni} = \frac{(0.250 \text{ g})}{(58.69 \text{ g/mol})}$$
$$= 0.004\,26 \text{ mol}$$

$$n_e = 0.004\,26 \text{ mol} \times \frac{2}{1}$$
$$= 0.008\,52 \text{ mol}$$

$$t = \frac{n_e F}{I}$$
$$= \frac{(0.008\,52 \text{ mol})\left(9.65 \times 10^4 \text{ C/mol}\right)}{(0.540 \text{ C/s})}$$
$$= 1.52 \times 10^3 \text{ s}$$

$$t = \left(1.52 \times 10^3 \text{ s}\right) \times \left(\frac{1 \text{ min}}{60 \text{ s}}\right)$$
$$= 25.4 \text{ min}$$

The time required is 25.4 min.

Section 14.4 11.

$Cr_2O_7^{2-}(aq) \rightarrow Cr(s)$
$Cr_2O_7^{2-}(aq) \rightarrow 2\ Cr(s)$
$Cr_2O_7^{2-}(aq) \rightarrow 2\ Cr(s) + 7\ H_2O(l)$
$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) \rightarrow 2\ Cr(s) + 7\ H_2O(l)$
$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 12\ e^- \rightarrow 2\ Cr(s) + 7\ H_2O(l)$

$$n_{Cr} = \frac{(17.8\ g)}{(52.00\ g/mol)}$$
$$= 0.342\ mol$$

$$n_e = 0.342\ mol \times \frac{12}{2}$$
$$= 2.05\ mol$$

$$t = 2.20\ h$$
$$= (2.20\ h) \times \left(\frac{3600\ s}{1\ h}\right)$$
$$= 7.92 \times 10^3\ s$$

$$I = \frac{n_e F}{t}$$
$$= \frac{(2.05\ mol)\left(9.65 \times 10^4\ C/mol\right)}{\left(7.92 \times 10^3\ s\right)}$$
$$= 25.0\ C/s$$

The average current required is 25.0 C/s.

Section 14.4 12.

Prediction

$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

$$t = (6.0 \text{ min}) \times \left(\frac{60 \text{ s}}{1 \text{ min}}\right)$$
$$= 360 \text{ s}$$

$$n_e = \frac{It}{F}$$
$$= \frac{(3.46 \text{ C/s})(360 \text{ s})}{(9.65 \times 10^4 \text{ C/mol})}$$
$$= 0.012\,91 \text{ mol}$$

$$n_{Sn} = 0.012\,91 \text{ mol} \times \frac{1}{2}$$
$$= 0.006\,45 \text{ mol}$$

$$m_{Sn} = (0.006\,45 \text{ mol})(118.71 \text{ g/mol})$$
$$= 0.766 \text{ g}$$

The mass of tin plated onto the can will be 0.766 g.

Analysis

m(Sn) deposited on can = 118.05 g – 117.34 g
= 0.71 g

The mass of tin electroplated onto the can was 0.71 g.

Evaluation

$$\% \text{ difference} = \frac{|(0.71 \text{ g} - 0.766 \text{ g})|}{0.766 \text{ g}} \times 100$$
$$= 7.3\%$$

The prediction is judged to be verified. A response that the result is inconclusive is also acceptable.

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